

# Chemistry 12

## Solutions Manual Part B

### Unit 3 Energy Changes and Rates of Reaction

#### Solutions to Practice Problems in Chapter 5 Energy Changes

##### Calculating the Absorption of Heat (Student textbook page 281)

1. How much heat must be added to a 36.2 g sample of iron to increase its temperature by 250.0°C?

##### What Is Required?

You have to calculate the amount of heat that must be added to a sample of iron to increase its temperature by 250.0°C.

##### What Is Given?

You know the mass of iron:  $m = 36.2$  g

You know the change in the temperature:  $\Delta T = 250.0^\circ\text{C}$

You have the specific heat capacity of iron from **Table 5.1** on page 280 of the student text:  $c_{\text{Fe}} = 0.449$  J/g•°C

Plan Your Strategy	Act on Your Strategy
Substitute the given information into the formula $Q = mc\Delta T$ . Solve for $Q$ .	$Q = mc\Delta T$ $= (36.2 \text{ g}) \left( 0.449 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \right) (250.0^\circ\text{C})$ $= 4063.45 \text{ J}$ $= 4.06 \text{ kJ}$ <p>The iron must gain 4060 J or 4.06 kJ of heat energy.</p>

##### Check Your Solution

Check that the information has been substituted correctly. The answer has the correct number of significant digits.

2. How much heat must be added to 128.62 g of steam at 126.0°C to increase its temperature to 189.5°C?

### What Is Required?

You must determine the amount of heat that must be added to a sample of steam to increase its temperature from 126.0°C to 189.5°C.

### What Is Given?

You know the mass of steam:  $m = 128.62 \text{ g}$

You know the initial temperature:  $T_{\text{initial}} = 126.0^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 189.5^\circ\text{C}$

You have the specific heat capacity of steam from Table 5.1 on page 280 of the student textbook:  $c_{\text{H}_2\text{O(g)}} = 2.02 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Determine the temperature change of the steam.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 189.5^\circ\text{C} - 126.0^\circ\text{C}$ $= 63.5^\circ\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the quantity of heat.	$Q = mc\Delta T$ $= (128.6 \cancel{\text{g}}) \left( 2.02 \frac{\text{J}}{\cancel{\text{g}} \cdot \cancel{^\circ\text{C}}} \right) (63.5 \cancel{^\circ\text{C}})$ $= 16\,495.5 \text{ J}$ $= 16.5 \text{ kJ}$ <p>The steam must gain 16500 J or 16.5 kJ of heat energy.</p>

### Check Your Solution

Check that the information has been substituted correctly. The answer has the correct number of significant digits.

3. A  $1.00 \times 10^2$  g sample of ethanol at  $25.0^\circ\text{C}$  is heated until it reaches  $50.0^\circ\text{C}$ . How much thermal energy does the ethanol gain?

### What Is Required?

You must calculate the thermal energy gained by a sample of ethanol.

### What Is Given?

You know the mass of ethanol:  $m = 1.00 \times 10^2$  g

You know the initial temperature:  $T_{\text{initial}} = 25.0^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 50.0^\circ\text{C}$

You have the specific heat capacity of ethanol from Table 5.1 on page 280 of the student textbook:  $c_{\text{ethanol}} = 2.44 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Determine the temperature change of the ethanol.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 50.0^\circ\text{C} - 25.0^\circ\text{C}$ $= 25.0^\circ\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the quantity of heat gained.	$Q = mc\Delta T$ $= (1 \times 10^2 \cancel{\text{g}}) \left( 2.44 \frac{\text{J}}{\cancel{\text{g}} \cdot \cancel{^\circ\text{C}}} \right) (25.0 \cancel{^\circ\text{C}})$ $= 6.10 \times 10^3 \text{ J}$ $= 6.10 \text{ kJ}$ <p>The ethanol gains <math>6.10 \times 10^3</math> J or 6.10 kJ of heat energy.</p>

### Check Your Solution

Check that the information has been substituted correctly. The answer has the correct number of significant digits.

4. Beaker A contains 50 g of liquid at room temperature. The beaker is heated until the liquid increases in temperature by 10°C. Beaker B contains 100 g of the same liquid at room temperature. The beaker is also heated until the liquid increases in temperature by 10°C. In which beaker does the liquid absorb more heat? Explain your answer.

**What Is Required?**

You must compare the heat absorbed by two samples of the same liquid.

**What Is Given?**

You know the mass of sample A is 50 g and that the mass of sample B is 100 g. Both liquids increase in temperature by 10°C.

Plan Your Strategy	Act on Your Strategy
Use the formula $Q = mc\Delta T$ to compare the quantities of heat gained.	<p>The liquid in beaker B absorbs twice as much heat because it has twice the mass.</p> $Q_A = m_A c \Delta T$ $= (50 \text{ g}) c \Delta T$ $Q_B = m_B c \Delta T$ $= (100 \text{ g}) c \Delta T$ $= 2[(50 \text{ g}) c \Delta T]$ $= 2[Q_A]$ <p>Therefore, <math>Q_B = 2Q_A</math>.</p>

**Check Your Solution**

The answer is reasonable since  $Q$  and  $m$  are directly related in the equation  $Q = mc\Delta T$ .

5. How much heat is released when the temperature of 789 g of liquid ammonia decreases from 82.7°C to 25.0°C?

### What Is Required?

You must determine the quantity of heat released when a sample of liquid ammonia is cooled.

### What Is Given?

You know the mass of ammonia:  $m = 789 \text{ g}$

You know the initial temperature:  $T_{\text{initial}} = 82.7^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 25.0^\circ\text{C}$

You have the specific heat capacity of ammonia from **Table 5.1** on page 280 of the student textbook:  $c_{\text{NH}_3(\ell)} = 4.70 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Determine the change in temperature of the liquid ammonia.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 25.0^\circ\text{C} - 82.70^\circ\text{C}$ $= -57.7^\circ\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the quantity of heat gained.	$Q = mc\Delta T$ $= (789 \cancel{\text{g}}) \left( 4.70 \frac{\text{J}}{\cancel{\text{g}} \cdot \cancel{^\circ\text{C}}} \right) (-57.7 \cancel{^\circ\text{C}})$ $= -213\,968.9 \text{ J}$ $= -2.14 \times 10^2 \text{ kJ}$ <p>The quantity of heat released is 214 000 J or <math>2.14 \times 10^2 \text{ kJ}</math>.</p>

### Check Your Solution

The answer seems reasonable and has the correct number of significant digits. The negative sign indicates that heat energy was given off.

6. A solid substance has a mass of 250.00 g. It is cooled by 25.00°C and loses 4.937 kJ of heat. What is the specific heat capacity of the substance? Identify the substance using the values in **Table 5.1** (student textbook page 280).

### What Is Required?

You must determine the specific heat capacity of a sample of solid.

### What Is Given?

You know the mass of the solid:  $m = 250.00 \text{ g}$

You know the change in temperature of the solid:  $\Delta T = 25.00^\circ\text{C}$

You know the quantity of heat lost:  $Q = -4.937 \text{ kJ}$

Plan Your Strategy	Act on Your Strategy
Substitute the given information into the formula $Q = mc\Delta T$ . Then solve for $c$ .	$\Delta T = -25.00^\circ\text{C}$  $Q = mc\Delta T$ $-4.937 \text{ kJ} = (250.00 \text{ g}) \times c \times (-25.00^\circ\text{C})$ $-4937 \text{ J} = (250.00 \text{ g}) \times c \times (-25.00^\circ\text{C})$ $c = 0.7899 \text{ J/g}\cdot^\circ\text{C}$
Use the specific heat capacity information in <b>Table 5.1</b> to identify the solid.	0.79 J/g $\cdot^\circ\text{C}$ is the specific heat capacity of granite.

### Check Your Solution

The specific heat capacity is in the range listed for solids in **Table 5.1** and seems reasonable. The answer shows close agreement with the value listed in **Table 5.1**.

7. The specific heat capacity of a compound used in fireworks is  $0.800 \text{ J/g}\cdot^{\circ}\text{C}$ . If it takes  $8.04 \times 10^3 \text{ J}$  to heat this material from  $20.0^{\circ}\text{C}$  to  $925.0^{\circ}\text{C}$ , what mass of compound was used?

### What Is Required?

You must determine the mass of a sample that has absorbed a given amount of heat energy.

### What Is Given?

You know the specific heat capacity of the compound:  $c = 0.800 \text{ J/g}\cdot^{\circ}\text{C}$

You know the initial temperature:  $T_{\text{initial}} = 20.0^{\circ}\text{C}$

You know the final temperature:  $T_{\text{final}} = 925.0^{\circ}\text{C}$

You know the quantity of heat absorbed:  $Q = 8.04 \times 10^3 \text{ J}$

Plan Your Strategy	Act on Your Strategy
Determine the temperature change of the compound.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 925.0^{\circ}\text{C} - 20.0^{\circ}\text{C}$ $= 905^{\circ}\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the mass.	$Q = mc\Delta T$ $8.04 \times 10^3 \text{ J} = m \left( 0.800 \frac{\text{J}}{\text{g}\cdot^{\circ}\text{C}} \right) (905.0^{\circ}\text{C})$ $m = \frac{8.04 \times 10^3 \cancel{\text{J}}}{\left( 0.800 \frac{\cancel{\text{J}}}{\text{g}\cdot^{\circ}\text{C}} \right) (905.0^{\circ}\text{C})}$ $m = 11.1 \text{ g}$ <p>The mass of the compound is 11.1 g.</p>

### Check Your Solution

The mass seems reasonable for a firecracker and the answer has the correct number of significant digits.

8. One litre of water at  $1.00^{\circ}\text{C}$  is warmed by the addition of  $4.00\text{ kJ}$  of heat energy. What is the final temperature of the water? (1.00 L of water has a mass of 1.00 kg)

### What Is Required?

You must find the final temperature of a sample of water that has gained heat energy.

### What Is Given?

You know the mass of water:  $m = 1.00\text{ kg}$ , or  $1 \times 10^3\text{ g}$

You know the initial temperature:  $T_{\text{initial}} = 1.00^{\circ}\text{C}$

You know the amount of heat absorbed:  $Q = 4.00\text{ kJ}$ , or  $4000\text{ J}$

You have the specific heat capacity of water from Table 5.1 on page 280 of the student textbook:  $c_{\text{H}_2\text{O}(\ell)} = 4.19\text{ J/g}\cdot^{\circ}\text{C}$

Plan Your Strategy	Act on Your Strategy
Use the formula $Q = mc\Delta T$ to determine the change in temperature, $\Delta T$ .	$Q = mc\Delta T$ $4000\text{ J} = (1.00 \times 10^3\text{ g}) \left( 4.19 \frac{\text{J}}{\text{g}\cdot^{\circ}\text{C}} \right) \Delta T$ $\Delta T = 0.9546^{\circ}\text{C}$
Calculate the final temperature, $T_{\text{final}}$ .	$0.9546^{\circ}\text{C} = T_{\text{final}} - T_{\text{initial}}$ $= T_{\text{final}} - 1.00^{\circ}\text{C}$ $T_{\text{final}} = 1.95^{\circ}\text{C}$ <p>The final temperature of the water is <math>1.95^{\circ}\text{C}</math>.</p>

### Check Your Solution

The answer is reasonable for this amount of heat energy added to 1.00 kg of water. The answer has the correct number of significant digits.



9. On a warm day, how much solar energy does a 3.982 kg piece of concrete absorb as heat energy if its temperature increases from 13.60°C to 14.50°C?

### What Is Required?

You need to calculate the quantity of heat absorbed by a piece of concrete.

### What Is Given?

You know the mass of concrete:  $m = 3.982 \text{ kg}$

You know the initial temperature:  $T_{\text{initial}} = 13.60^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 14.50^\circ\text{C}$

You have the specific heat capacity of concrete from **Table 5.1** on page 280 of the student textbook:  $c_{\text{concrete}} = 0.88 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Determine the change of temperature of the concrete.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 14.50^\circ\text{C} - 13.60^\circ\text{C}$ $= 0.90^\circ\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the quantity of heat absorbed.	$Q = mc\Delta T$ $= (3.982 \text{ kg}) \left( 0.88 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \right) (0.90^\circ\text{C})$ $= 3153.7 \text{ J}$ <p>The concrete absorbed 3154 J or 3.2 kJ of energy.</p>

### Check Your Solution

The answer is reasonable and shows the correct number of significant digits.

10. You have samples of the air and hydrogen gas at room temperature, both having a mass of 10.00 g.
- Compare the change in temperature of these two samples if each gains 500.0 J of thermal energy.
  - Suggest a reason for the difference in the temperature changes.

### What Is Required?

- You must calculate and compare the temperature change of samples of hydrogen and of air that have the same mass and gain the same amount of thermal energy.
- You must suggest a reason for the difference in temperature.

### What Is Given?

You know the mass of hydrogen:  $m_{\text{H}_2} = 10.00 \text{ g}$

You know the mass of the air:  $m_{\text{air}} = 10.00 \text{ g}$

You know the amount of heat absorbed by each sample:  $Q = 500 \text{ J}$

You have the specific heat capacities of hydrogen and air from **Table 5.1** on page 280 of the student textbook:  $c_{\text{H}_2} = 14.304 \text{ J/g}\cdot^\circ\text{C}$  and  $c_{\text{air}} = 1.01 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
<p><b>a.</b> Use the formula <math>Q = mc\Delta T</math> to calculate the change in temperature, <math>\Delta T</math>, for each sample.</p>	$Q = m_{\text{air}} c_{\text{air}} \Delta T_{\text{air}}$ $500 \cancel{\text{J}} = (10.00 \cancel{\text{g}}) \left( 1.01 \frac{\cancel{\text{J}}}{\cancel{\text{g}}\cdot^\circ\text{C}} \right) \Delta T_{\text{air}}$ $\Delta T_{\text{air}} = 49.5^\circ\text{C}$ $Q = m_{\text{H}} c_{\text{H}} \Delta T_{\text{H}}$ $500 \cancel{\text{J}} = (10.00 \cancel{\text{g}}) \left( 14.304 \frac{\cancel{\text{J}}}{\cancel{\text{g}}\cdot^\circ\text{C}} \right) \Delta T_{\text{H}}$ $\Delta T_{\text{H}} = 3.496^\circ\text{C}$ <p><math>\Delta T</math> for air is <math>49.5^\circ\text{C}</math> and <math>\Delta T</math> for hydrogen is <math>3.49^\circ\text{C}</math>; the change in temperature of the hydrogen gas is about 14 times less than the change in temperature of the air.</p>

**b.** The molecules that make up air, primarily oxygen,  $\text{O}_2(\text{g})$ , and nitrogen,  $\text{N}_2(\text{g})$ , have a greater mass than hydrogen molecules,  $\text{H}_2(\text{g})$ . There will be more molecules in 10.0 g of  $\text{H}_2(\text{g})$  than in 10.0 g of air. The absorbed energy will be distributed among a larger number of  $\text{H}_2(\text{g})$  molecules and each will gain less energy than in the air sample. Since temperature is a measure of the average kinetic energy of the particles in a sample, the  $\text{H}_2(\text{g})$  sample will experience a smaller temperature change as calculated. (Going further: the average molar mass of molecules in air is about 29 g/mol. The molar mass of hydrogen is 2 g/mol, about 14 times less.)

**Check Your Solution**

The answer is reasonable since the average molar mass of air is about 14 times the molar mass of hydrogen gas.

**Calculating an Enthalpy Change**  
**Using Enthalpy Data to Determine the Mass of Products**  
**(Student textbook page 299)**

11. Pentane reacts with an excess of oxygen to produce carbon dioxide and water vapour. What is the enthalpy change of the reaction per mole of each of the following gases? The enthalpy of combustion for pentane,  $\text{C}_5\text{H}_{12}(\ell)$ , is  $-3509.0 \text{ kJ/mol}$ .
- oxygen
  - carbon dioxide
  - water

**What Is Required?**

You need to calculate the enthalpy change per mole for  $\text{O}_2(\text{g})$ ,  $\text{CO}_2(\text{g})$ , and  $\text{H}_2\text{O}(\text{g})$  during the burning of pentane.

**What Is Given?**

You know that pentane is burned in an excess of oxygen, so complete combustion occurs.

You know that the enthalpy of combustion for pentane  $\text{C}_5\text{H}_{12}(\ell)$ :  $\Delta H_{\text{comb}}^\circ = -3509.0 \text{ kJ/mol}$ .

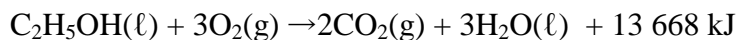
Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the complete combustion of pentane.	$\text{C}_5\text{H}_{12}(\ell) + 8\text{O}_2(\text{g}) \rightarrow 5\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{g})$ $\Delta H_{\text{comb}}^\circ = -3509.0 \text{ kJ/mol}$
a. Use the mole ratio in the balanced chemical equation to calculate the enthalpy change per mole of $\text{O}_2(\text{g})$ .	$\Delta H_r(\text{per mole O}_2) = \frac{n_{\text{C}_5\text{H}_{12}}}{n_{\text{O}_2}} \times \frac{\Delta H_{\text{comb}}^\circ}{n_{\text{C}_5\text{H}_{12}}}$ $= \frac{1 \cancel{\text{mol C}_5\text{H}_{12}}}{8 \text{ mol O}_2} \times \frac{-3509.0 \text{ kJ}}{1 \cancel{\text{mol C}_5\text{H}_{12}}}$ $= -438.62 \text{ kJ}$
b. Use the mole ratio in the balanced chemical equation to calculate the enthalpy change per mole of $\text{CO}_2(\text{g})$ .	$\Delta H_r(\text{per mole CO}_2) = \frac{n_{\text{C}_5\text{H}_{12}}}{n_{\text{CO}_2}} \times \frac{\Delta H_{\text{comb}}^\circ}{n_{\text{C}_5\text{H}_{12}}}$ $= \frac{1 \cancel{\text{mol C}_5\text{H}_{12}}}{5 \text{ mol CO}_2} \times \frac{-3509.0 \text{ kJ}}{1 \cancel{\text{mol C}_5\text{H}_{12}}}$ $= -701.80 \text{ kJ}$

<p>c. Use the mole ratio in the balanced chemical equation to calculate the enthalpy change per mole of H<sub>2</sub>O(g).</p>	$\begin{aligned}\Delta H_r(\text{per mole H}_2\text{O}) &= \frac{n_{\text{C}_5\text{H}_{12}}}{n_{\text{H}_2\text{O}}} \times \frac{\Delta H_{\text{comb}}^\circ}{n_{\text{C}_5\text{H}_{12}}} \\ &= \frac{1 \cancel{\text{mol C}_5\text{H}_{12}}}{6 \text{ mol H}_2\text{O}} \times \frac{-3509.0 \text{ kJ}}{1 \cancel{\text{mol C}_5\text{H}_{12}}} \\ &= -584.83 \text{ kJ}\end{aligned}$
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### Check Your Solution

The enthalpy per mol is inversely related to the number of moles. This is a reasonable result. The number of significant digits in each answer agrees with the given data.

12. What is the enthalpy change when 4.608 g of ethanol,  $\text{C}_2\text{H}_5\text{OH}(\ell)$ , undergoes complete combustion?



### What Is Required?

You need to determine the enthalpy change when a given mass of ethanol undergoes complete combustion.

### What Is Given?

You know the mass of ethanol:  $m_{\text{ethanol}} = 4.608 \text{ g}$

You know the balanced chemical equation for the reaction:



Plan Your Strategy	Act on Your Strategy
Determine the amount in moles, $n$ , of ethanol by using its mass and molar mass. Use the formula $n = \frac{m}{M}$ .	$n_{\text{C}_2\text{H}_5\text{OH}} = \frac{m}{M}$ $= \frac{4.608 \cancel{\text{g}}}{46.08 \cancel{\text{g}}/\text{mol}}$ $= 0.1000 \text{ mol}$
Use the formula $\Delta H = n\Delta H_{\text{comb}}^\circ$ to determine the enthalpy change.	$\Delta H = n\Delta H_{\text{comb}}^\circ$ $(0.1000 \cancel{\text{mol}})(-1366.8 \text{ kJ}/\cancel{\text{mol}})$ $= -136.7 \text{ kJ}$

### Check Your Solution

The given mass is less than the mass of 1 mol of  $\text{C}_2\text{H}_5\text{OH}(\ell)$  and the enthalpy change is less than the enthalpy change per mol of  $\text{C}_2\text{H}_5\text{OH}(\ell)$ . This is a reasonable result and shows the correct number of significant digits.

13. Determine the thermal energy released by the combustion of each of the following samples of hydrocarbons to  $\text{CO}_2(\text{g})$  and  $\text{H}_2\text{O}(\text{l})$  in an open system.

- a. 56.78 g of hexane,  $\text{C}_6\text{H}_{14}(\text{l})$
- b. 1.36 kg of octane,  $\text{C}_8\text{H}_{18}(\text{l})$
- c.  $2.344 \times 10^4$  g of heptane,  $\text{C}_7\text{H}_{16}(\text{l})$

#### What Is Required?

You need to calculate the thermal energy released during combustion of each sample of gas.

#### What Is Given?

You know the mass of hexane,  $\text{C}_6\text{H}_{14}(\text{l})$ :  $m = 56.78$  g

You know the mass of octane,  $\text{C}_8\text{H}_{18}(\text{l})$ :  $m = 1.36$  kg

You know the mass of heptane,  $\text{C}_7\text{H}_{16}(\text{l})$ :  $m = 2.344 \times 10^4$  g

From **Table 5.4** (student textbook page 295), you know the heat of combustion for each gas:

$\Delta H^\circ_{\text{comb}}$  for hexane =  $-4163.2$  kJ/mol

$\Delta H^\circ_{\text{comb}}$  for octane =  $-5470.3$  kJ/mol

$\Delta H^\circ_{\text{comb}}$  for heptane =  $-4817.0$  kJ/mol

Plan Your Strategy	Act on Your Strategy
Use the molar mass of each gas and the formula $n = \frac{m}{M}$ to calculate the amount in moles of each sample.	<p>a. <math>n_{\text{C}_6\text{H}_{14}} = \frac{m}{M}</math></p> $= \frac{56.78 \cancel{\text{g}}}{86.20 \cancel{\text{g}}/\text{mol}}$ $= 0.65870 \text{ mol}$ <p>b. <math>n_{\text{C}_8\text{H}_{18}} = \frac{m}{M}</math></p> $= \frac{1.36 \times 10^3 \cancel{\text{g}}}{114.26 \cancel{\text{g}}/\text{mol}}$ $= 11.9026 \text{ mol}$ <p>c. <math>n_{\text{C}_7\text{H}_{16}} = \frac{m}{M}</math></p> $= \frac{2.344 \times 10^4 \cancel{\text{g}}}{100.23 \cancel{\text{g}}/\text{mol}}$ $= 233.8621 \text{ mol}$

<p>Use the formula <math>\Delta H_r = n\Delta H_{\text{comb}}^\circ</math> to calculate the thermal energy released by the combustion of each sample.</p>	<p><b>a.</b> <math>\Delta H_r = n_{\text{C}_6\text{H}_{14}} \Delta H_{\text{comb}}^\circ</math>  <math>= (0.65870 \text{ mol})(-4163.2 \text{ kJ/mol})</math>  <math>= -2742.17 \text{ kJ}</math>  <math>= -2742 \text{ kJ}</math></p> <p><b>b.</b> <math>\Delta H_r = n_{\text{C}_8\text{H}_{18}} \Delta H_{\text{comb}}^\circ</math>  <math>= (11.9026 \text{ mol})(-5470.3 \text{ kJ/mol})</math>  <math>= -65\,110.792 \text{ kJ}</math>  <math>= -6.51 \times 10^4 \text{ kJ}</math></p> <p><b>c.</b> <math>\Delta H_r = n_{\text{C}_7\text{H}_{16}} \Delta H_{\text{comb}}^\circ</math>  <math>= (233.8611 \text{ mol})(-4817.0 \text{ kJ/mol})</math>  <math>= -1\,126\,513.8 \text{ kJ}</math>  <math>= -1.126 \times 10^6 \text{ kJ}</math></p>
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### Check Your Solution

Check that the given data has been substituted correctly. The answers seem reasonable and show the correct number of significant digits.



14. What is the enthalpy change for the combustion of a 1.00 g sample of methane,  $\text{CH}_4(\text{g})$ , under standard conditions?

### What Is Required?

You need to determine the enthalpy change when a given mass of methane,  $\text{CH}_4(\text{g})$ , undergoes combustion.

### What Is Given?

You know the mass of methane,  $\text{CH}_4(\text{g})$ :  $m = 1.00 \text{ g}$

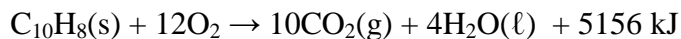
From **Table 5.4** (student textbook page 295), you know the heat of combustion for methane:  $\Delta H_{\text{comb}}^\circ = -890.8 \text{ kJ/mol}$

Plan Your Strategy	Act on Your Strategy
Use the molar mass of methane and the formula $n = \frac{m}{M}$ to calculate the amount in moles, $n$ , of $\text{CH}_4(\text{g})$ .	$n_{\text{CH}_4} = \frac{m}{M}$ $= \frac{1.00 \cancel{\text{g}}}{16.05 \cancel{\text{g}}/\text{mol}}$ $= 0.062305 \text{ mol}$
Use the formula $\Delta H_{\text{r}} = n\Delta H_{\text{comb}}^\circ$ to calculate the thermal energy released by the combustion of the sample.	$\Delta H_{\text{r}} = n_{\text{CH}_4} \Delta H_{\text{comb}}^\circ$ $= (0.062305 \cancel{\text{mol}})(-890.8 \text{ kJ}/\cancel{\text{mol}})$ $= -55.5 \text{ kJ}$

### Check Your Solution

The enthalpy change would be expected to be small for a small mass of methane. The answer is reasonable and shows the correct number of significant digits.

15. Naphthalene,  $\text{C}_{10}\text{H}_8(\text{s})$ , is an organic compound used in the manufacture of dyes, plastics, and insecticides, especially insecticides that combat moths. What mass of this compound will release 500.0 kJ of thermal energy?



### What Is Required?

You need to calculate the mass of naphthalene,  $\text{C}_{10}\text{H}_8(\text{s})$ , that can be combusted to release a given amount of thermal energy.

### What Is Given?

You know the balanced thermochemical equation for the combustion of naphthalene,  $\text{C}_{10}\text{H}_8(\text{s})$ .

Plan Your Strategy	Act on Your Strategy
Use the mole ratio in the balanced thermochemical equation to determine the amount in moles, $n$ , of $\text{C}_{10}\text{H}_8(\text{s})$ that were combusted to produce 500.0 kJ of thermal energy.	$\frac{1 \text{ mol C}_{10}\text{H}_8}{5156 \text{ kJ}} = \frac{n_{\text{C}_{10}\text{H}_8}}{500.0 \text{ kJ}}$ $n_{\text{C}_{10}\text{H}_8} = \frac{(1 \text{ mol C}_{10}\text{H}_8)(500.0 \cancel{\text{kJ}})}{5156 \cancel{\text{kJ}}}$ $= 0.096974 \text{ mol}$
Use the molar mass of naphthalene and the formula $m = n \times M$ to determine the mass, $m$ , of $\text{C}_{10}\text{H}_8(\text{s})$ .	$m_{\text{C}_{10}\text{H}_8} = n \times M$ $= 0.096974 \cancel{\text{mol}} \times 128.18 \text{ g}/\cancel{\text{mol}}$ $= 12.43 \text{ g}$

### Check Your Solution

The mass of naphthalene seems reasonable and the answer shows the correct number of significant digits.

16. What mass, in kilograms, of methanol must be combusted to generate  $5.39 \times 10^5$  kJ of thermal energy?

### What Is Required?

You need to determine the mass, in kilograms, of methanol that must be combusted to produce a given amount of thermal energy.

### What Is Given?

You know the amount of energy produced:  $\Delta H_r = -5.39 \times 10^5$  kJ

You know that the fuel to be burned is methanol.

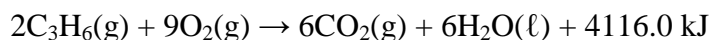
From **Table 5.4** (student textbook page 295), you know that the heat of combustion of methanol  $\Delta H_{\text{comb}}^\circ = -726.1$  kJ/mol.

Plan Your Strategy	Act on Your Strategy
Use the heat of combustion, $\Delta H_{\text{comb}}^\circ$ , of methanol, $\text{CH}_3\text{OH}(\ell)$ , and the formula $\Delta H_r = n\Delta H_{\text{comb}}^\circ$ to determine the amount in moles of $\text{CH}_3\text{OH}(\ell)$ that was combusted.	$\Delta H_r = n\Delta H_{\text{comb}}^\circ$ $-5.39 \times 10^5 \text{ kJ} = n(-726.1 \text{ kJ/mol})$ $n = \frac{-5.39 \times 10^5 \cancel{\text{kJ}}}{-726.1 \cancel{\text{kJ}}/\text{mol}}$ $= 742.32 \text{ mol}$
Convert the amount in moles of $\text{CH}_3\text{OH}(\ell)$ to mass, $m$ , using the molar mass of methanol and the formula $m = n \times M$ .	$m_{\text{CH}_3\text{OH}} = n \times M$ $= 742.32 \cancel{\text{mol}} \times 32.05 \text{ g}/\cancel{\text{mol}}$ $= 23\,791 \text{ g}$ $= 23.8 \text{ kg}$

### Check Your Solution

The mass of methanol seems reasonable for this large amount of energy to be given off during combustion. The number of significant digits in the answer is correct.

17. Answer the questions for the combustion of propene,  $\text{C}_3\text{H}_6(\text{g})$ , using the following equation:



- Determine the standard molar enthalpy of combustion of propene.
- What is the maximum amount of thermal energy available from the combustion of 5.00 g of  $\text{C}_3\text{H}_6(\text{g})$ ?

### What Is Required?

- You need to determine the standard molar enthalpy of combustion for propene  $\text{C}_3\text{H}_6(\text{g})$ :
- You need to calculate the amount of thermal energy released when a given mass of  $\text{C}_3\text{H}_6(\text{g})$  burns.

### What Is Given?

- You know the balanced thermochemical equation for the combustion of 2 mol of  $\text{C}_3\text{H}_6(\text{g})$ .
- You know the mass of propene:  $m_{\text{C}_3\text{H}_6} = 5.00 \text{ g}$

Plan Your Strategy	Act on Your Strategy
a. Rewrite the thermochemical equation for the combustion for 1 mol of $\text{C}_3\text{H}_6(\text{g})$ . The quantity of heat shown in this equation is the standard heat of combustion.	$\text{C}_3\text{H}_6(\text{g}) + \frac{9}{2}\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\ell) + 2058.0 \text{ kJ}$ $\Delta H_{\text{comb}}^{\circ} = -2058.0 \text{ kJ/mol}$
b. Use the molar mass of propene and the formula $n = \frac{m}{M}$ to calculate the amount in moles, $n$ , of $\text{C}_3\text{H}_6(\text{g})$ .	$n = \frac{m}{M}$ $n_{\text{C}_3\text{H}_6} = \frac{5.00 \cancel{\text{g}}}{42.09 \cancel{\text{g}}/\text{mol}}$ $= 0.11879 \text{ mol}$
Use the formula $\Delta H_{\text{r}} = n\Delta H_{\text{comb}}^{\circ}$ to calculate the amount of heat given off.	$\Delta H_{\text{r}} = n\Delta H_{\text{comb}}^{\circ}$ $= (0.11879 \cancel{\text{mol}})(-2058.0 \text{ kJ}/\cancel{\text{mol}}) =$ $= 244.469 \text{ kJ}$ <p>The thermal energy released is 244 kJ.</p>

### Check Your Solution

- The answer has the correct number of significant digits.
- The answer seems reasonable since the mass burned is less than 1 mol. The correct number of significant digits is shown.

18. When 0.050 mol of a hydrocarbon undergoes complete combustion, 110.95 kJ of heat is given off.
- What is the enthalpy of combustion for this hydrocarbon?
  - Refer to **Table 5.4** (student textbook page 295) and identify the hydrocarbon.

### What Is Required?

- You need to determine the enthalpy of combustion of a hydrocarbon.
- You need to identify the hydrocarbon.

### What Is Given?

You know the amount in moles of hydrocarbon:  $n = 0.050$  mol

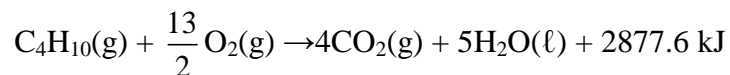
You know the amount of heat energy given off:  $Q = -110.95$  kJ

Plan Your Strategy	Act on Your Strategy
a. Calculate the amount of energy given off from the complete combustion of 1 mol of hydrocarbon.	<p>a. <math display="block">\frac{0.050 \text{ mol}}{-110.95 \text{ kJ}} = \frac{1}{\Delta H_{\text{comb}}^{\circ}}</math></p> $\Delta H_{\text{comb}}^{\circ} = \frac{-110.95 \text{ kJ}}{0.050 \text{ mol}}$ $= -2219 \text{ kJ/mol}$ $= -2.2 \times 10^3 \text{ kJ/mol}$
b. Refer to <b>Table 5.4</b> .	b. From <b>Table 5.4</b> , the hydrocarbon is propane.

### Check Your Solution

The answer is reasonable for 1 mol and shows the correct number of significant digits.

19. What mass of butane,  $\text{C}_4\text{H}_{10}(\text{g})$ , must undergo complete combustion to produce 1.00 MJ of heat according to the following equation?



### What Is Required?

You need to calculate the mass of butane,  $\text{C}_4\text{H}_{10}(\text{g})$ , that must be burned to produce a given amount of heat.

### What Is Given?

You know amount of heat energy produced:  $\Delta H = 1.00 \text{ MJ}$

From the balanced thermochemical equation, you know the heat of combustion per mol of butane,  $\text{C}_4\text{H}_{10}(\text{g})$ .

Plan Your Strategy	Act on Your Strategy
Convert 1.00 MJ to kJ.	$1.00 \text{ MJ} = 1.00 \times 10^3 \text{ kJ}$
Use the mole ratio in the balanced thermochemical equation to calculate the amount in moles, $n$ , of butane, $\text{C}_4\text{H}_{10}(\text{g})$ , that is combusted.	$\frac{-2877.6 \text{ kJ}}{1 \text{ mol C}_4\text{H}_{10}} = \frac{-1.00 \times 10^3 \text{ kJ}}{n_{\text{C}_4\text{H}_{10}}}$ $n_{\text{C}_4\text{H}_{10}} = \frac{(1 \text{ mol C}_4\text{H}_{10})(-1.00 \times 10^3 \cancel{\text{kJ}})}{-2877.6 \cancel{\text{kJ}}}$ $= 0.34751 \text{ mol}$
Convert the amount in moles of $\text{C}_4\text{H}_{10}(\text{g})$ to mass, $m$ , using the molar mass of butane and the formula $m = n \times M$ .	$m_{\text{C}_4\text{H}_{10}} = n \times M$ $= 0.34751 \cancel{\text{mol}} \times 58.14 \text{ g}/\cancel{\text{mol}}$ $= 20.2 \text{ g}$

### Check Your Solution

Since the given amount of heat energy is less than the heat of combustion per mol. the mass burned should be less than the mass of 1 mol. the answer is reasonable. The answer has the correct number of significant digits.

20. The standard molar heat of combustion of propan-2-ol,  $\text{CH}_3\text{CH}(\text{OH})\text{CH}_3(\ell)$ , is  $-2006 \text{ kJ/mol}$ .
- Write the thermochemical equation for the combustion of 1 mol of this compound to form carbon dioxide and water.
  - How much heat is given off when 25.00 g of water is produced?

### What Is Required?

- You need to determine the amount of heat given off when a given mass of water is produced during the burning of propan-2-ol,  $\text{CH}_3\text{CH}(\text{OH})\text{CH}_3(\ell)$ .
- You need to determine the amount of heat given off when 25.00 g of water is produced.

### What Is Given?

You know standard heat of combustion of  $\text{CH}_3\text{CH}(\text{OH})\text{CH}_3(\ell)$ :  $\Delta H_{\text{comb}}^\circ = -2006 \text{ kJ/mol}$

You know the chemical formula for propan-2-ol,  $\text{CH}_3\text{CH}(\text{OH})\text{CH}_3(\ell)$ .

You know that the products of complete combustion are  $\text{CO}_2(\text{g})$  and  $\text{H}_2\text{O}(\ell)$ .

Plan Your Strategy	Act on Your Strategy
<b>a.</b> Write a balanced chemical equation showing propanol-2-ol and oxygen, $\text{O}_2(\text{g})$ , as reactants and carbon dioxide, $\text{CO}_2(\text{g})$ , and water, $\text{H}_2\text{O}(\ell)$ , as products. Divide each coefficient by a factor so that 1 mol of propan-2-ol is shown in the balanced chemical equation. Since the standard molar heat of combustion is a negative number, you know the reaction is exothermic.	$\text{CH}_3\text{CH}(\text{OH})\text{CH}_3(\ell) + \frac{9}{2} \text{O}_2(\text{g}) \rightarrow$ $3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\ell) + 2006 \text{ kJ}$
<b>b.</b> Use the molar mass of water and the formula $n = \frac{m}{M}$ to determine the amount in moles, $n$ , of $\text{H}_2\text{O}(\ell)$ .	$n = \frac{m}{M}$ $n_{\text{H}_2\text{O}} = \frac{25.00 \text{ g}}{18.02 \text{ g/mol}}$ $= 1.387 \text{ 34 mol}$
Use the mole ratio in the balanced chemical equation to calculate the amount of heat given off.	$\frac{-2006 \text{ kJ}}{4 \text{ mol H}_2\text{O}(\ell)} = \frac{x}{1.387 \text{ 34 mol H}_2\text{O}(\ell)}$ $x = -695.8 \text{ kJ}$

**Check Your Solution**

**a.** The balanced equation shows 1 mol of propan-2-ol. The heat term is on the right side of the equation since a combustion reaction is exothermic.

**b.** The amount of heat given off is about  $\frac{1}{4}$  the heat of combustion for

$\text{CH}_3\text{CH}(\text{OH})\text{CH}_3(\ell)$ . Since the balanced thermochemical equation shows 4 mol of  $\text{H}_2\text{O}(\ell)$  produced, the answer is reasonable and shows the correct number of significant digits.



## Determining the Enthalpy Change of a Reaction (Student textbook page 305)

21. A pellet of potassium hydroxide, KOH(s), having a mass of 0.648 g, is dissolved in 40.0 mL of water in an insulated cup. The temperature of the water increases from 22.6°C to 27.8°C. What is the molar enthalpy of solution,  $\Delta H_{\text{solution}}$ , for KOH(s)? Assume that the solution has a density and a specific heat capacity equal to that of water.

### What Is Required?

You need to calculate the molar enthalpy of solution,  $\Delta H_{\text{solution}}$ , for KOH(s).

### What Is Given?

You know the mass of the pellet of KOH(s):  $m_{\text{KOH(s)}} = 0.648 \text{ g}$

You know the mass of the solution, KOH(aq):  $m_{\text{solution}} = 40.0 \text{ g}$

You know the initial temperature:  $T_{\text{initial}} = 22.6^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 27.8^\circ\text{C}$

You know the specific heat capacity of  $\text{H}_2\text{O}(\ell)$ :  $4.19 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Convert the mass of KOH(s) to an amount in moles, $n$ , using the molar mass of potassium hydroxide and the formula $n = \frac{m}{M}$ .	$n = \frac{m}{M}$ $n_{\text{KOH(s)}} = \frac{0.648 \cancel{\text{g}}}{56.11 \cancel{\text{g}}/\text{mol}}$ $= 0.011\,548\,7 \text{ mol}$
Determine the temperature change of the system.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 27.8^\circ\text{C} - 22.6^\circ\text{C}$ $= 5.2^\circ\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the amount of heat absorbed by the solution.	$Q = m_{\text{solution}} c_{\text{solution}} \Delta T_{\text{solution}}$ $= (40.0 \cancel{\text{g}})(4.19 \text{ J}/\cancel{\text{g}\cdot^\circ\text{C}})(5.2 \cancel{^\circ\text{C}})$ $= 871.52 \text{ J}$ $= 0.871\,52 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find $\Delta H$ , the change in the thermal energy of the system.	$\Delta H = -Q$ $= -0.871\,52 \text{ kJ}$

Use the formula $\Delta H_{\text{solution}} = \frac{\Delta H}{n}$ to determine the molar enthalpy of solution.	$\Delta H_{\text{solution}} = \frac{\Delta H}{n}$ $= \frac{-0.871\ 52\ \text{kJ}}{0.011\ 548\ 7\ \text{mol}}$ $= -75.5\ \text{kJ/mol}$
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### Check Your Solution

Since much less than 1 mol of KOH(s) was dissolved, it is reasonable that a heat of solution per mol will be much more than the heat absorbed by the 40.0 g of solution. You know the reaction is exothermic since the temperature increases. The calculated value of  $\Delta H$  is negative. The answer shows the correct number of significant digits.

22. When 5.022 g of sodium hydrogen carbonate,  $\text{NaHCO}_3(\text{s})$ , reacts completely with 80.00 mL of acetic acid,  $\text{CH}_3\text{COOH}(\text{aq})$ , the temperature increases from  $18.6^\circ\text{C}$  to  $28.4^\circ\text{C}$ .



Assume that the acid solution has the same density and specific heat capacity as water and that the mass of the final solution is 80.00 g. Calculate the molar enthalpy of reaction,  $\Delta H_{\text{r}}$ .

### What Is Required?

You need to calculate the molar enthalpy of reaction.

### What Is Given?

You know the mass of  $\text{NaHCO}_3(\text{s})$ :  $m_{\text{NaHCO}_3} = 5.022 \text{ g}$

You know the mass of  $\text{CH}_3\text{COOH}(\text{aq})$  solution:  $m_{\text{solution}} = 80.0 \text{ g}$

You know the initial temperature:  $T_{\text{initial}} = 18.6^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 28.4^\circ\text{C}$

You know the specific heat capacity of  $\text{H}_2\text{O}(\text{l})$ :  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Convert the mass of $\text{NaHCO}_3(\text{s})$ to amount in moles, $n$ , using the molar mass of sodium hydrogen carbonate and the formula $n = \frac{m}{M}$ .	$n = \frac{m}{M}$ $n_{\text{NaHCO}_3(\text{s})} = \frac{5.022 \cancel{\text{g}}}{84.01 \cancel{\text{g}}/\text{mol}}$ $= 0.059\,778 \text{ mol}$
Determine the temperature change of the system.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 28.4^\circ\text{C} - 18.6^\circ\text{C}$ $= 9.8^\circ\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the amount of heat absorbed by the solution.	$Q = m_{\text{solution}} c_{\text{solution}} \Delta T_{\text{solution}}$ $= (80.0 \cancel{\text{g}})(4.19 \text{ J}/\cancel{\text{g}}\cdot^\circ\text{C})(9.8\cancel{^\circ\text{C}})$ $= 3284.96 \text{ J}$ $= 3.284\,96 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find $\Delta H$ , the change in the thermal energy of the system.	$\Delta H = -Q$ $= -3.284\,96 \text{ kJ}$

<p>Use the formula <math>\Delta H_r = \frac{\Delta H}{n}</math> to determine the molar enthalpy of reaction.</p>	$\Delta H_r = \frac{\Delta H}{n}$ $= \frac{-3.284\ 96\ \text{kJ}}{0.0597\ 78\ \text{mol}}$ $= 54.952\ 65\ \text{kJ/mol}$ <p>The enthalpy of reaction is <math>-55.0\ \text{kJ/mol}</math>.</p>
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### Check Your Solution

The answer is reasonable for this amount of  $\text{NaHCO}_3(\text{s})$  reacting. You know the reaction is exothermic since the temperature increases. The calculated value of  $\Delta H$  is negative. The answer has the correct number of significant digits.

23. Sodium reacts violently to form sodium hydroxide when placed in water, as shown in the following equation:  $2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + \text{H}_2\text{(g)}$

Determine an experimental value for the molar enthalpy of reaction for sodium given the following data:

mass of sodium, Na(s): 0.37 g

mass of water in calorimeter: 175 g

initial temperature of water: 19.30°C

final temperature of mixture: 25.70°C

### What Is Required?

You need to calculate the molar enthalpy of reaction.

### What Is Given?

You know the mass of sodium, Na(s):  $m_{\text{Na}} = 0.37 \text{ g}$

You know the mass of water in the calorimeter:  $m_{\text{calorimeter}} = 175 \text{ g}$

You know the initial temperature:  $T_{\text{initial}} = 19.30^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 25.70^\circ\text{C}$

You know the specific heat capacity of  $\text{H}_2\text{O(l)}$ :  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^\circ\text{C}$

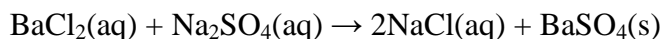
Plan Your Strategy	Act on Your Strategy
Convert the mass of Na(s) to amount in moles, $n$ , using the molar mass of sodium and the formula $n = \frac{m}{M}$ .	$n = \frac{m}{M}$ $n_{\text{Na}} = \frac{0.37 \cancel{\text{g}}}{22.99 \cancel{\text{g}}/\text{mol}}$ $= 0.016\,093 \text{ mol}$
Determine the temperature change of the system.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 25.70^\circ\text{C} - 19.30^\circ\text{C}$ $= 6.40^\circ\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the amount of heat absorbed by the solution.	$Q = m_{\text{solution}} c_{\text{solution}} \Delta T_{\text{solution}}$ $= (175 \cancel{\text{g}})(4.19 \text{ J}/\cancel{\text{g}}\cdot^\circ\text{C})(6.40^\circ\text{C})$ $= 4692.8 \text{ J}$ $= 4.692\,8 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find $\Delta H$ , the change in the thermal energy of the system.	$\Delta H = -Q$ $= -4.692\,8 \text{ kJ}$

<p>Use the formula <math>\Delta H_r = \frac{\Delta H}{n}</math> to determine the molar enthalpy of reaction.</p>	$\Delta H_r = \frac{\Delta H}{n}$ $= \frac{-4.6928 \text{ kJ}}{0.016093 \text{ mol}}$ $= -291.605 \text{ kJ/mol}$ $= -2.9 \times 10^2 \text{ kJ/mol}$ <p>The molar enthalpy of reaction is <math>-2.9 \times 10^2</math> kJ/mol of sodium.</p>
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### Check Your Solution

The answer is reasonable for this amount of Na(s) reacting. You know the reaction is exothermic since the temperature increases. The calculated value of  $\Delta H$  is negative. The answer has the correct number of significant digits.

24. In a simple calorimeter, 250.0 mL of 0.120 mol/L barium chloride,  $\text{BaCl}_2(\text{aq})$ , is mixed with 150.0 mL of 0.200 mol/L sodium sulfate,  $\text{Na}_2\text{SO}_4(\text{aq})$ . A precipitate of barium sulfate,  $\text{BaSO}_4(\text{s})$ , forms. The initial temperature of the two solutions is  $20.00^\circ\text{C}$ . After mixing, the final temperature of the solutions is  $20.49^\circ\text{C}$ . Calculate the enthalpy of reaction, in kJ/mol, of  $\text{Na}_2\text{SO}_4(\text{aq})$ . Assume that the solutions have densities and specific heat capacities equivalent to those of water.



### What Is Required?

You need to calculate the enthalpy of reaction.

### What Is Given?

You know the balanced chemical equation for the reaction that occurs in aqueous solution.

You know the concentration of  $\text{BaCl}_2(\text{aq})$ :  $c_{\text{BaCl}_2} = 0.120 \text{ mol/L}$

You know the volume of  $\text{BaCl}_2(\text{aq})$ :  $V_{\text{BaCl}_2} = 250.0 \text{ mL}$

You know the mass of  $\text{BaCl}_2(\text{aq})$ :  $m_{\text{BaCl}_2} = 250.0 \text{ g}$

You know the concentration of  $\text{Na}_2\text{SO}_4(\text{aq})$ :  $c_{\text{Na}_2\text{SO}_4} = 0.200 \text{ mol/L}$

You know the volume of  $\text{Na}_2\text{SO}_4(\text{aq})$ :  $V_{\text{Na}_2\text{SO}_4} = 150.0 \text{ mL}$

You know the mass of  $\text{Na}_2\text{SO}_4(\text{aq})$ :  $m_{\text{Na}_2\text{SO}_4} = 150.0 \text{ g}$

You know the total mass of the solutions:  $m_{\text{T}} = 400.0 \text{ g}$

You know the initial temperature:  $T_{\text{initial}} = 20.0^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 20.49^\circ\text{C}$

You know the specific heat capacity of each solution is the same as the specific heat capacity of  $\text{H}_2\text{O}(\ell)$ :  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Use the formula $n = cV$ to calculate the amount in moles of each reactant.	$n = cV$ $n_{\text{BaCl}_2} = (0.120 \text{ mol/L})(0.250 \text{ L})$ $= 0.300 \text{ mol}$ $n = cV$ $n_{\text{Na}_2\text{SO}_4} = (0.200 \text{ mol/L})(0.150 \text{ L})$ $= 0.300 \text{ mol}$
Determine if there is a limiting reactant.	Neither reactant is limiting.
Determine the temperature change of the system.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 20.49^\circ\text{C} - 20.00^\circ\text{C}$ $= 0.49^\circ\text{C}$

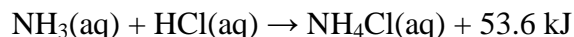
Use the formula $Q = mc\Delta T$ to calculate the heat absorbed by the mixture.	$Q = m_{\text{mixture}} c_{\text{mixture}} \Delta T_{\text{mixture}}$ $= (400 \text{ g})(4.19 \text{ J/g}\cdot^{\circ}\text{C})(0.49^{\circ}\text{C})$ $= 821.24 \text{ J}$ $= 8.2124 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find $\Delta H$ , the change in the thermal energy of the system.	$\Delta H = -Q$ $= -0.821 \text{ 24 kJ}$
Use the formula $\Delta H_{\text{r}} = \frac{\Delta H}{n}$ to determine the molar heat of reaction.	$\Delta H_{\text{r}} = \frac{\Delta H}{n}$ $= \frac{-0.821 \text{ 24 kJ}}{0.0300 \text{ mol}}$ $= -27.370 \text{ 8 kJ/mol}$ $= -27.4 \text{ kJ/mol Na}_2\text{SO}_4(\text{aq})$ <p>The enthalpy of reaction is <math>-27.4 \text{ kJ/mol}</math>.</p>

### Check Your Solution

You know the reaction is exothermic since the temperature increases. The calculated value of  $\Delta H$  is negative. The answer has the correct number of significant digits and seems reasonable.



25. A neutralization reaction occurs when 100.0 mL of 0.200 mol/L aqueous ammonia,  $\text{NH}_3(\text{aq})$ , and 200.0 mL of 0.200 mol/L hydrochloric acid,  $\text{HCl}(\text{aq})$ , are mixed in an insulated cup.



Assuming that the two solutions have the same density and specific heat capacity as water, what temperature change is expected after mixing?

### What Is Required?

You need to calculate the temperature change for the reaction.

### What Is Given?

You know the balanced thermochemical equation for the reaction that occurs in aqueous solution.

From this, you know the heat of reaction:  $\Delta H_r = -53.6 \text{ kJ/mol}$

You know the concentration of  $\text{NH}_3(\text{aq})$  solution:  $c_{\text{NH}_3} = 0.200 \text{ mol/L}$

You know the volume of  $\text{NH}_3(\text{aq})$  solution:  $V_{\text{NH}_3} = 100.0 \text{ mL}$

You know the mass of  $\text{NH}_3(\text{aq})$  solution:  $m_{\text{NH}_3} = 100.0 \text{ g}$

You know the concentration of  $\text{HCl}(\text{aq})$  solution:  $c_{\text{HCl}} = 0.200 \text{ mol/L}$

You know the volume of  $\text{HCl}(\text{aq})$  solution:  $V_{\text{HCl}} = 250.0 \text{ mL}$

You know the mass of  $\text{HCl}(\text{aq})$  solution:  $m_{\text{HCl}} = 250.0 \text{ g}$

You know the total mass of solution:  $m_T = 350.0 \text{ g}$

You know the specific heat capacity of each solution is the same as the specific heat capacity of  $\text{H}_2\text{O}(\ell)$ :  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^\circ\text{C}$

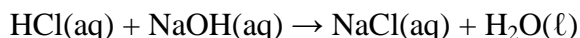
Plan Your Strategy	Act on Your Strategy
Use the formula $n = cV$ to calculate the amount in moles of each reactant.	$n = cV$ $n_{\text{NH}_3} = (0.200 \text{ mol/L})(0.100 \text{ L})$ $= 0.0200 \text{ mol}$ $n = cV$ $n_{\text{HCl}} = (0.200 \text{ mol/L})(0.200 \text{ L})$ $= 0.0400 \text{ mol}$
Determine if there is a limiting reactant.	The balanced chemical equation indicates that $\text{NH}_3(\text{aq})$ and $\text{HCl}(\text{aq})$ react in a mole ratio of 1:1. By inspection, since there is a lesser amount in moles of $\text{NH}_3(\text{aq})$ , this reactant is limiting. The amount in moles of $\text{NH}_3(\text{aq})$ must be used to calculate $\Delta H_r$ .

Use the formula $\Delta H = n\Delta H_r$ to find $\Delta H_r$ .	$\Delta H = n\Delta H_r$ $\Delta H_r = \frac{n_{\text{NH}_3}}{\Delta H}$ $= \frac{0.020 \text{ mol}}{-53.6 \text{ kJ/mol}}$ $= -1.072 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $\Delta H$ to find $Q$ , the amount of heat absorbed by the solutions.	$Q = -\Delta H$ $= 1072 \text{ J}$
Use the formula $Q = mc\Delta T$ to calculate the temperature change, $\Delta T$ .	$Q = mc\Delta T$ $1072 \text{ J} = (300 \text{ g})(4.19 \text{ J/g} \cdot ^\circ\text{C})(\Delta T)$ $\Delta T = \frac{(300 \cancel{\text{g}})(4.19 \cancel{\text{J}} / \cancel{\text{g}} \cdot ^\circ\text{C})}{1072 \cancel{\text{J}}}$ $= 0.852 \text{ } ^\circ\text{C}$ <p>The temperature increases by <math>0.853^\circ\text{C}</math>.</p>

### Check Your Solution

You know the reaction is exothermic so the temperature must increase. This seems to be a reasonable temperature change. The answer has the correct number of significant digits.

26. In a simple calorimeter, 150.0 mL of 1.000 mol/L NaOH(aq) is mixed with 150.0 mL of 1.000 mol/L HCl(aq). If both solutions were initially at 25.00°C and after mixing the temperature increased to 30.00°C, what is the enthalpy of reaction as written? Assume that the solutions have a density of 1.000 g/mL and a specific heat capacity of 4.19 J/g•°C.



### What Is Required?

You need to calculate the enthalpy of reaction.

### What Is Given?

You know the balanced chemical equation for the reaction that occurs in aqueous solution.

You know the concentration of NaOH(aq) solution:  $c_{\text{NaOH}} = 1.000 \text{ mol/L}$

You know the volume of NaOH(aq) solution:  $V_{\text{NaOH}} = 150.0 \text{ mL}$

You know the mass of NaOH(aq) solution:  $m_{\text{NaOH}} = 150.0 \text{ g}$

You know the concentration of HCl(aq) solution:  $c_{\text{HCl}} = 1.000 \text{ mol/L}$

You know the volume of HCl(aq) solution:  $V_{\text{HCl}} = 150.0 \text{ mL}$

You know the mass of HCl(aq) solution:  $m_{\text{HCl}} = 150.0 \text{ g}$

You know the total mass of the solutions: 300.0 g

You know the initial temperature:  $T_{\text{initial}} = 25.00^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 30.00^\circ\text{C}$

You know the specific heat capacity of each solution is the same as the specific heat capacity of  $\text{H}_2\text{O(l)}$ :  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Use the formula $n = cV$ to calculate the amount in moles of each reactant.	$n = cV$ $n_{\text{HCl}} = (1.000 \text{ mol/L})(0.1500 \text{ L})$ $= 0.1500 \text{ mol}$
Determine the temperature change of the system.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 30.00^\circ\text{C} - 25.00^\circ\text{C}$ $= 5.00^\circ\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the heat absorbed by the mixture.	$Q = m_{\text{mixture}} c_{\text{mixture}} \Delta T_{\text{mixture}}$ $= (300 \text{ g})(4.19 \text{ J/g}\cdot^\circ\text{C})(5.00^\circ\text{C})$ $= 6285 \text{ J}$ $= 6.285 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find $\Delta H$ , the change in the thermal energy of the system.	$\Delta H = -Q$ $= -6.285 \text{ kJ}$

<p>Use the formula <math>\Delta H_r = \frac{\Delta H}{n}</math> to determine the molar heat of reaction.</p>	$\Delta H_r^\circ = \frac{-6.285 \text{ kJ}}{0.150 \text{ 00 mol}}$ $= -41.9 \text{ kJ/mol}$ <p>The enthalpy of reaction is <math>-41.9 \text{ kJ/mol}</math>.</p>
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### Check Your Solution

You know the reaction is exothermic since the temperature increases. The calculated value of  $\Delta H$  is negative. The answer has the correct number of significant digits and seems reasonable.

27. The enthalpy of solution for sodium hydroxide, NaOH(s), is given as  $-55.0 \text{ kJ/mol}$ . A chemist prepares  $250.0 \text{ mL}$  of a  $0.100 \text{ mol/L}$  solution of NaOH(aq). Assuming that this solution has the same specific heat capacity and density as water, by how much should the water temperature increase as the NaOH(s) dissolves?

### What Is Required?

You need to determine the temperature change as NaOH(s) dissolves.

### What Is Given?

You know the enthalpy of solution:  $-55.0 \text{ kJ/mol}$

You know the concentration of NaOH(aq) solution:  $c_{\text{NaOH}} = 0.100 \text{ mol/L}$

You know the volume NaOH(aq) solution:  $V_{\text{NaOH}} = 100.0 \text{ mL}$

You know the mass of NaOH(aq) solution:  $m_{\text{NaOH}} = 100.0 \text{ g}$

You know the specific heat capacity of the solution is the same as the specific heat capacity of  $\text{H}_2\text{O}(\ell)$ :  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Use the formula $n = cV$ to calculate the amount in moles, $n$ , of NaOH(aq).	$n = cV$ $n_{\text{NaOH}} = (0.100 \text{ mol/L})(0.2500 \text{ L})$ $= 0.02500 \text{ mol}$
Use the formula $\Delta H_{\text{solution}} = \frac{\Delta H}{n}$ to determine the amount of heat given off during dissolving.	$\Delta H_{\text{solution}} = \frac{\Delta H}{n}$ $-55.0 \text{ kJ/mol} = \frac{\Delta H}{0.02500 \text{ mol}}$ $\Delta H = (-55.0 \text{ kJ/mol})(0.02500 \text{ mol})$ $= -1.375 \text{ kJ}$ $= -1375 \text{ J}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $\Delta H$ to find the amount of heat, $Q$ , absorbed by the surroundings.	$Q_{\text{surroundings}} = -\Delta H$ $= 1375 \text{ J}$
Use the formula $Q = mc\Delta T$ to calculate the change in temperature, $\Delta T$ .	$Q = mc\Delta T$ $1375 \text{ J} = (250 \text{ g})(4.19 \text{ J/g}\cdot^\circ\text{C})(\Delta T)$ $\Delta T = \frac{1375 \text{ J}}{(250 \text{ g})(4.19 \text{ J/g}\cdot^\circ\text{C})}$ $= 1.31^\circ\text{C}$ <p>The temperature of the water increases by <math>1.31^\circ\text{C}</math>.</p>

### Check Your Solution

You know the reaction is exothermic, so the temperature must increase. The change in temperature seems reasonable. The answer has the correct number of significant digits.

28. A neutralization reaction occurs when 120.00 mL of 0.500 mol/L LiOH and 160.00 mL of 0.375 mol/L HNO<sub>3</sub>(aq) are mixed in an insulated cup. Initially, the solutions are at the same temperature. If the highest temperature reached during mixing was 24.5°C, what was the initial temperature of the solutions?



Assume that both of these solutions have a density of 1.00 g/mL and a specific heat capacity of 4.19 J/g •°C.

### What Is Required?

You need to determine the initial temperature of solutions used in a neutralization reaction.

### What Is Given?

You know the concentration,  $c$ , of LiOH(aq) solution:  $c_{\text{LiOH}} = 0.500 \text{ mol/L}$

You know the volume of LiOH(aq) solution:  $V_{\text{LiOH}} = 120.0 \text{ mL}$

You know the mass of LiOH(aq) solution:  $m_{\text{LiOH}} = 120.0 \text{ g}$

You know the concentration,  $c$ , of HNO<sub>3</sub>(aq) solution:  $c_{\text{HNO}_3} = 0.3750 \text{ mol/L}$

You know the volume of HNO<sub>3</sub>(aq) solution:  $V_{\text{HNO}_3} = 160.0 \text{ mL}$

You know the mass of HNO<sub>3</sub>(aq) solution:  $m_{\text{HNO}_3} = 160.0 \text{ g}$

You know the total mass of the solutions:  $m_{\text{T}} = 280.0 \text{ g}$

You know the final temperature:  $T_{\text{final}} = 24.5^\circ\text{C}$

You know the specific heat capacity of each solution is the same as the specific heat capacity of H<sub>2</sub>O(l):  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g} \cdot ^\circ\text{C}$

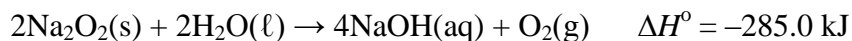
Plan Your Strategy	Act on Your Strategy
Use the formula $n = cV$ to calculate the amount in moles, $n$ , of each reactant.	$n = cV$ $n_{\text{LiOH}} = (0.500 \text{ mol/L})(0.1200 \text{ L})$ $= 0.0600 \text{ mol}$ $n = cV$ $n_{\text{HNO}_3} = (0.375 \text{ mol/L})(0.1600 \text{ L})$ $= 0.0600 \text{ mol}$
Determine if there is a limiting reactant.	Neither reactant is limiting.

Use the formula $\Delta H_{\text{solution}} = \frac{\Delta H}{n}$ to determine the amount of heat given off in the neutralization.	$\Delta H_{\text{solution}} = \frac{\Delta H}{n}$ $-53.1 \text{ kJ/mol} = \frac{\Delta H}{0.0600 \text{ mol}}$ $\Delta H = (-53.1 \text{ kJ/mol})(0.0600 \text{ mol})$ $= -3.186 \text{ kJ}$ $= -3186 \text{ J}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $\Delta H$ to find the amount of heat, $Q$ , absorbed by the surroundings.	$Q_{\text{neutralization}} = -\Delta H$ $= 3186 \text{ J}$
Use the formula $Q = mc\Delta T$ to calculate the change in temperature, $\Delta T$ .	$Q = mc\Delta T$ $3186 \text{ J} = (280 \text{ g})(4.19 \text{ J/g} \cdot ^\circ\text{C})(\Delta T)$ $\Delta T = \frac{3186 \cancel{\text{J}}}{(280 \cancel{\text{g}})(4.19 \cancel{\text{J}}/\cancel{\text{g}} \cdot ^\circ\text{C})}$ $= 2.7^\circ\text{C}$
Use the formula $\Delta T = T_{\text{final}} - T_{\text{initial}}$ to determine the initial temperature of the solutions.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $T_{\text{initial}} = T_{\text{final}} - \Delta T$ $= 24.5^\circ\text{C} - 2.7^\circ\text{C}$ $= 21.8^\circ\text{C}$ <p>The initial temperature of the solutions was <math>21.8^\circ\text{C}</math>.</p>

### Check Your Solution

You know the reaction is exothermic. Since the temperature must increase, the initial temperature must be lower than the final temperature. This seems to be a reasonable initial temperature. The answer has the correct number of significant digits.

29. Peroxides will react to release oxygen when added to water. By how much would the water temperature change if 7.800 g of sodium peroxide,  $\text{Na}_2\text{O}_2(\text{s})$ , is added to 110.00 mL of water?



### What Is Required?

You need to determine the change in temperature when sodium peroxide,  $\text{Na}_2\text{O}_2(\text{s})$ , is dissolved in water.

### What Is Given?

You know the balanced chemical equation for the reaction and the enthalpy of reaction:  $\Delta H^\circ = -285.0 \text{ kJ}$

You know the volume of water:  $V_{\text{H}_2\text{O}} = 110.0 \text{ mL}$

You know the mass,  $m$ , of water:  $m_{\text{H}_2\text{O}} = 110.0 \text{ g}$

You know the specific heat capacity of  $\text{H}_2\text{O}(\ell)$ :  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^\circ\text{C}$

You know the mass,  $m$ , of  $\text{Na}_2\text{O}_2(\text{s})$ :  $m_{\text{Na}_2\text{O}_2(\text{s})} = 7.800 \text{ g}$

Plan Your Strategy	Act on Your Strategy
Convert the mass of $\text{Na}_2\text{O}_2(\text{s})$ to amount in moles using the molar mass, $M$ , of $\text{Na}_2\text{O}_2(\text{s})$ and the formula $n = \frac{m}{M}$ .	$n = \frac{m}{M}$ $n_{\text{Na}_2\text{O}_2} = \frac{7.800 \text{ g}}{77.98 \text{ g/mol}}$ $= 0.1000 \text{ mol}$
Let $x$ represent $\Delta H_r$ . Use the mole ratio in the balanced chemical equation to calculate the enthalpy change.	$\frac{2 \text{ mol Na}_2\text{O}_2}{-285.0 \text{ kJ}} = \frac{0.1000 \text{ mol Na}_2\text{O}_2}{x}$ $x = \frac{(0.1000 \text{ mol Na}_2\text{O}_2)(285.0 \text{ kJ})}{2 \text{ mol Na}_2\text{O}_2}$ $= -14.25 \text{ kJ}$ $\Delta H_r = -14.25 \text{ kJ}$ $= -14\,250 \text{ J}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $\Delta H_r$ to find the amount of heat, $Q$ , absorbed by the surroundings.	$Q = -\Delta H_r$ $= 14\,250 \text{ J}$

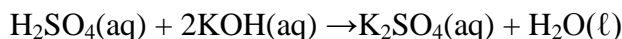


Use the formula $Q = mc\Delta T$ to calculate the change in temperature, $\Delta T$ .	$Q = mc\Delta T$ $14\,250\text{ J} = (110.0\text{ g})(4.19\text{ J/g}\cdot^{\circ}\text{C})(\Delta T)$ $\Delta T = \frac{14\,250\cancel{\text{ J}}}{(110.0\cancel{\text{ g}})(4.19\cancel{\text{ J}}/\cancel{\text{ g}}\cdot^{\circ}\text{C})}$ $= 30.9^{\circ}\text{C}$ <p>The temperature of the water increases by <math>30.9^{\circ}\text{C}</math>.</p>
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### Check Your Solution

You know the reaction is exothermic so the temperature must increase. This seems to be a reasonable temperature change. The answer has the correct number of significant digits.

30. In an insulated calorimeter, 200.0 mL of 1.00 mol/L potassium hydroxide, KOH(aq), is mixed with an equal volume of 1.00 mol/L sulfuric acid, H<sub>2</sub>SO<sub>4</sub>(aq). The temperature increases by 6.50°C. Assume that the solutions have the same density and specific heat capacity as water.



What is the enthalpy of neutralization?

### What Is Required?

You need to calculate the molar enthalpy of neutralization.

### What Is Given?

You know the balanced chemical equation for the reaction that occurs in aqueous solution.

You know the concentration of KOH(aq) solution:  $c_{\text{KOH}} = 1.000 \text{ mol/L}$

You know the volume of KOH(aq) solution:  $V_{\text{KOH}} = 200.0 \text{ mL}$

You know the mass of KOH(aq) solution:  $m_{\text{KOH}} = 200.0 \text{ g}$

You know the concentration of H<sub>2</sub>SO<sub>4</sub>(aq) solution:  $c_{\text{H}_2\text{SO}_4} = 1.000 \text{ mol/L}$

You know the volume, of H<sub>2</sub>SO<sub>4</sub>(aq) solution:  $V_{\text{H}_2\text{SO}_4} = 200.0 \text{ mL}$

You know the mass,  $m$ , of H<sub>2</sub>SO<sub>4</sub>(aq) solution:  $m_{\text{H}_2\text{SO}_4} = 200.0 \text{ g}$

You know the total mass,  $m$ , of solution:  $m_{\text{T}} = 400.0 \text{ g}$

You know the change in temperature,  $\Delta T$ : 6.50°C

You know the specific heat capacity of solutions is that same as the specific heat capacity of H<sub>2</sub>O(ℓ):  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Use the formula $n = cV$ to calculate the amount in moles, $n$ , of each reactant.	$n = cV$ $n_{\text{H}_2\text{SO}_4} = (1.00 \text{ mol/L})(0.200 \text{ L})$ $= 0.200 \text{ mol}$ $n = cV$ $n_{\text{KOH}} = (1.00 \text{ mol/L})(0.200 \text{ L})$ $= 0.200 \text{ mol}$

Determine which reagent is limiting.	<p>From the balanced equation, the mole ratio <math>\text{H}_2\text{SO}_4(\text{aq}):\text{KOH}(\text{aq}) = 1:2</math></p> $\frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol KOH}} = \frac{x}{0.200 \text{ mol KOH}}$ $x = 0.100 \text{ mol H}_2\text{SO}_4(\text{aq})$ <p>The <math>\text{H}_2\text{SO}_4(\text{aq})</math> is in excess. The calculation of the enthalpy of neutralization must be based on the amount in moles of the limiting reactant <math>\text{KOH}(\text{aq})</math>.</p>
Use the formula $Q = mc\Delta T$ to calculate the heat absorbed by the surroundings.	$Q = m_{\text{solution}} c_{\text{solution}} \Delta T_{\text{solution}}$ $= (400 \text{ g})(4.19 \text{ J/g}\cdot^\circ\text{C})(6.50^\circ\text{C})$ $= 10\,984 \text{ J}$ $= 10.984 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find the amount of heat, $\Delta H$ , given off during the reaction.	$\Delta H = -Q$ $= -10.984 \text{ kJ}$
Use the formula $\Delta H_{\text{r}} = \frac{\Delta H}{n}$ and the amount in moles of $\text{KOH}(\text{aq})$ to determine the molar heat of reaction.	$\Delta H_{\text{neutralization}} = \frac{\Delta H}{n_{\text{KOH}}}$ $= \frac{-10.894 \text{ kJ}}{0.200 \text{ mol}}$ $= -54.47 \text{ kJ/mol}$ <p>The enthalpy of neutralization is <math>-54.47 \text{ kJ/mol}</math>.</p>

### Check Your Solution

You know the reaction is exothermic since the temperature increased. The calculated value of  $\Delta H$  is negative. The answer has the correct number of significant digits and seems reasonable.

## Calculating Thermal Energy in a Bomb Calorimeter (Student textbook page 309)

31. Predict the final temperature of a  $5.00 \times 10^2$  g iron ring that is initially at  $25.0^\circ\text{C}$  and is heated by combusting 4.95 g of ethanol,  $\text{C}_2\text{H}_5\text{OH}(\ell)$ , in an open system;  $\Delta H_{\text{comb}}$  for ethanol is  $-1234.8$  kJ/mol. (Note: These are not standard conditions.)

### What Is Required?

You need to find the final temperature of an iron ring heated by burning alcohol.

### What Is Given?

You know the mass of iron:  $m_{\text{Fe}} = 5.00 \times 10^2$  g

You know the initial temperature of the iron:  $T_{\text{initial}} = 25.0^\circ\text{C}$

You know the mass of ethanol burned:  $m_{\text{C}_2\text{H}_5\text{OH}} = 4.95$  g

You know the heat of combustion for ethanol:  $\Delta H_{\text{comb}}^\circ = -1234.8$  kJ/mol

Plan Your Strategy	Act on Your Strategy
Determine the amount in moles, $n$ , of $\text{C}_2\text{H}_5\text{OH}(\ell)$ combusted using the mass and molar mass of ethanol and the formula $n = \frac{m}{M}$ .	$n_{\text{C}_2\text{H}_5\text{OH}} = \frac{m}{M}$ $= \frac{4.95 \cancel{\text{g}}}{46.08 \cancel{\text{g}}/\text{mol}} = 0.1074 \text{ mol}$ $= 0.1074 \text{ mol}$
Use the formula $\Delta H = n\Delta H_{\text{comb}}$ to calculate the heat given off during combustion.	$\Delta H = n\Delta H_{\text{comb}}^\circ$ $= (0.1074 \cancel{\text{mol}})(-1234.8 \text{ kJ}/\cancel{\text{mol}})$ $= -132.617 \text{ kJ}$ $= -132\,617 \text{ J}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $\Delta H$ to find the amount of heat, $Q$ , absorbed by the iron.	$Q = -\Delta H$ $= 132\,617 \text{ J}$
Use the formula $Q = mc\Delta T$ to calculate the change in temperature, $\Delta T$ , of the iron.	$Q = mc\Delta T$ $132\,617 \text{ J} = (5.00 \text{ g})(0.449 \text{ J/g}\cdot^\circ\text{C})(\Delta T)$ $\Delta T = \frac{132\,617 \cancel{\text{J}}}{(5.00 \cancel{\text{g}})(0.449 \cancel{\text{J}}/\cancel{\text{g}}\cdot^\circ\text{C})}$ $= 590.7^\circ\text{C}$

Use  $\Delta T = T_{\text{final}} - T_{\text{initial}}$  to determine the final temperature.

$$\begin{aligned}\Delta T &= T_{\text{final}} - T_{\text{initial}} \\ T_{\text{final}} &= \Delta T + T_{\text{initial}} \quad \text{The final} \\ &= 590.7^{\circ}\text{C} + 25.00^{\circ}\text{C} \\ &= 615.7^{\circ}\text{C} \\ \text{temperature of the iron is } &616^{\circ}\text{C}.\end{aligned}$$

### Check Your Solution

The final temperature must be greater than the initial temperature since heat is gained by the iron. The answer is reasonable since a great deal of heat energy was absorbed and iron has a relatively low specific heat capacity. The final temperature has the correct number of significant digits.

32. Calculate the molar enthalpy of combustion of octane if 0.53 g of the fuel increased the temperature of a coffee can calorimeter (13 g of aluminum and  $2.50 \times 10^2$  mL of water) by  $17.2^\circ\text{C}$ . Remember to include the heat gained by not only the water but also by the aluminum can.

### What Is Required?

You need to calculate the molar enthalpy of combustion for octane.

### What Is Given?

You know the mass of octane:  $m_{\text{C}_8\text{H}_{18}} = 0.53 \text{ g}$

You know the mass of the aluminum can:  $m_{\text{Al}} = 13 \text{ g}$

You know the mass of water:  $m_{\text{H}_2\text{O}} = 2.50 \times 10^2 \text{ g}$

You know the change in the temperature:  $\Delta T = 17.2^\circ\text{C}$

You know the specific heat capacity of Al(s):  $c_{\text{Al}} = 0.897 \text{ J/g}\cdot^\circ\text{C}$

You know the specific heat capacity of solutions is the same as the specific heat capacity of  $\text{H}_2\text{O}(\ell)$ :  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Use the formula $Q = mc\Delta T$ to calculate the heat gained by the aluminum and the water.	$Q = (mc\Delta T)_{\text{Al}} + (mc\Delta T)_{\text{H}_2\text{O}}$ $= (13 \cancel{\text{g}})(0.897 \text{ J}/\cancel{\text{g}}\cdot^\circ\text{C})(17.2^\circ\cancel{\text{C}})$ $+ (2.50 \times 10^2 \cancel{\text{g}})(4.19 \text{ J}/\cancel{\text{g}}\cdot^\circ\text{C})(17.2^\circ\cancel{\text{C}})$ $= 200.569 \text{ J} + 18\,017 \text{ J}$ $= 18\,217.569 \text{ J}$ $= 18.217 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find the amount of heat, $\Delta H$ , given off during the combustion.	$\Delta H = -Q$ $= -18.217 \text{ kJ}$
Determine the amount in moles, $n$ , of $\text{C}_8\text{H}_{18}(\ell)$ combusted using the mass and molar mass of octane and the formula $n = \frac{m}{M}$ .	$n = \frac{m}{M}$ $n_{\text{C}_8\text{H}_{18}} = \frac{0.53 \cancel{\text{g}}}{114.26 \cancel{\text{g}}/\text{mol}}$ $= 0.004\,638 \text{ mol}$

<p>Use the formula <math>\Delta H_{\text{comb}} = \frac{\Delta H}{n}</math> to calculate the thermal energy released.</p>	$\Delta H_{\text{comb}} = \frac{\Delta H}{n}$ $\Delta H_{\text{comb}} = \frac{-18.217 \text{ kJ}}{0.004\,638 \text{ mol}}$ $= -3927.8 \text{ kJ/mol}$ <p>The enthalpy of combustion of octane is <math>-3.9 \times 10^3 \text{ kJ/mol}</math>.</p>
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### Check Your Solution

The enthalpy of combustion is negative as expected since heat is given off. The result is lower than the value in **Table 5.4** (student textbook page 295) because in this experimental situation, heat would be lost to the surroundings. The answer has the correct number of significant digits.

33. How much propane (in grams) would have to be combusted in an open system to increase the temperature of  $3.00 \times 10^2$  mL of water from  $20.00^\circ\text{C}$  to its boiling point? (The molar enthalpy of combustion of propane can be found in **Table 5.4** on page 295 of the student textbook.)

### What Is Required?

You need to determine the mass of propane that will increase the temperature of a sample of water to its boiling point.

### What Is Given?

You know the volume of water:  $V_{\text{H}_2\text{O}} = 3.00 \times 10^2$  mL

You know the mass of water:  $m_{\text{H}_2\text{O}} = 3.00 \times 10^2$  g

You know the initial temperature:  $T_{\text{initial}} = 20.00^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 100.00^\circ\text{C}$

You know the chemical formula for propane(g):  $\text{C}_3\text{H}_8$

You know the molar enthalpy of combustion for  $\text{C}_3\text{H}_8(\text{g})$ :  $\Delta H_{\text{comb}}^\circ = -2219.2$  kJ/mol

You know the specific heat capacity of  $\text{H}_2\text{O}(\ell)$ :  $c_{\text{H}_2\text{O}} = 4.19$  J/g $\cdot^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Use the formula $\Delta T = T_{\text{final}} - T_{\text{initial}}$ to determine the change in temperature.	$\begin{aligned}\Delta T &= T_{\text{final}} - T_{\text{initial}} \\ &= 100.00^\circ\text{C} - 20.00^\circ\text{C} \\ &= 80.00^\circ\text{C}\end{aligned}$
Use the formula $Q = mc\Delta T$ to calculate amount of heat, $Q$ , absorbed by the water.	$\begin{aligned}Q &= mc\Delta T \\ &= (3.00 \times 10^2 \text{ g})(4.19 \text{ J/g}\cdot^\circ\text{C})(80.00^\circ\text{C}) \\ &= 100\,560 \text{ J} \\ &= 100.560 \text{ kJ}\end{aligned}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to determine the amount of heat, $\Delta H$ , given off during the combustion.	$\begin{aligned}\Delta H &= -Q \\ &= -100.560 \text{ kJ}\end{aligned}$
Use the formula $\Delta H_{\text{comb}}^\circ = \frac{\Delta H}{n}$ to determine the amount in moles, $n$ , of $\text{C}_3\text{H}_8(\text{g})$ .	$\begin{aligned}\Delta H_{\text{comb}}^\circ &= \frac{\Delta H}{n} \\ -2219.2 \text{ kJ/mol} &= \frac{-100.560 \text{ kJ}}{n_{\text{C}_3\text{H}_8}} \\ n_{\text{C}_3\text{H}_8} &= \frac{-100.560 \text{ kJ}}{-2219.2 \text{ kJ/mol}} \\ &= 0.045\,31 \text{ mol}\end{aligned}$



Determine the mass, $m$ , of $\text{C}_3\text{H}_8(\text{g})$ required, using the molar mass and amount in moles of propane, and the formula $m = nM$ .	$m = nM$ $m_{\text{C}_3\text{H}_8} = (0.043\,31\cancel{\text{mol}})(44.11\,\text{g}/\cancel{\text{mol}})$ $= 1.9987\,\text{g}$ $= 2.00\,\text{g}$
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### Check Your Solution

The heat of combustion of  $\text{C}_3\text{H}_8(\text{g})$  is high. It is reasonable to expect a small mass of propane to change the temperature of water by  $80.00^\circ\text{C}$ . The answer has the correct number of significant digits.

34. A lab technician places a 5.00 g food sample into a bomb calorimeter with a heat capacity of 9.23 kJ/°C. The initial temperature of the calorimeter system is 21.0°C. After burning the food, the final temperature of the system is 32.0°C. How much thermal energy is released by the combustion of the food in kilojoules per gram?

### What Is Required?

You need to calculate the thermal energy released by the burning of a food sample.

### What Is Given?

You know the mass of food:  $m_{\text{food}} = 5.00 \text{ g}$

You know the heat capacity of the calorimeter:  $C = 9.23 \text{ kJ/}^\circ\text{C}$

You know the initial temperature:  $T_{\text{initial}} = 21.0^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 32.0^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Use $\Delta T = T_{\text{final}} - T_{\text{initial}}$ to determine the change in temperature.	$\begin{aligned}\Delta T &= T_{\text{final}} - T_{\text{initial}} \\ &= 32.0^\circ\text{C} - 21.0^\circ\text{C} \\ &= 11.0^\circ\text{C}\end{aligned}$
Use the formula $Q = C\Delta T$ to determine the amount of heat, $Q$ , gained by the calorimeter.	$\begin{aligned}Q &= C\Delta T \\ &= (9.23 \text{ kJ/}^\circ\text{C})(11.0^\circ\text{C}) \\ &= 101.53 \text{ kJ}\end{aligned}$
heat given off per gram of food = $\frac{Q}{m}$	$\begin{aligned}\text{heat given off} &= \frac{Q}{m} \\ &= \frac{101.53 \text{ kJ}}{5.00 \text{ g}} \\ &= 20.306 \text{ kJ/g}\end{aligned}$ <p>The heat given off when the food is burned is +20.3 kJ/g.</p>

### Check Your Solution

The quantity of heat is a positive value since it is absorbed by the calorimeter. This is a reasonable answer. The answer has the correct number of significant digits.

35. Determine the enthalpy of combustion of an unknown fuel if a 2.75 g sample increased the temperature of  $5.00 \times 10^2$  mL of hot chocolate ( $c = 3.75 \text{ J/g}\cdot^\circ\text{C}$ ) in a  $1.50 \times 10^2$  g glass mug ( $c = 0.84 \text{ J/g}\cdot^\circ\text{C}$ ) from  $10.00^\circ\text{C}$  to  $45.00^\circ\text{C}$ . Express the value for enthalpy of combustion in appropriate units.

### What Is Required?

You need to calculate the enthalpy of combustion of a fuel.

### What Is Given?

You know the mass,  $m$ , of fuel:  $m_{\text{fuel}} = 2.75 \text{ g}$

You know the volume,  $V$ , of hot chocolate:  $V_{\text{hot chocolate}} = 5.00 \times 10^2 \text{ mL}$

You know the mass,  $m$ , of hot chocolate:  $m_{\text{hot chocolate}} = 5.00 \times 10^2 \text{ g}$

You know the specific heat capacity of hot chocolate:  $c_{\text{hot chocolate}} = 3.75 \text{ J/g}\cdot^\circ\text{C}$

You know the mass of the glass mug:  $m_{\text{mug}} = 1.50 \times 10^2 \text{ g}$

You know the specific heat capacity of the glass mug:  $c_{\text{glass mug}} = 0.84 \text{ J/g}\cdot^\circ\text{C}$

You know the initial temperature:  $T_{\text{initial}} = 10.00^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 45.00^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Use $\Delta T = T_{\text{final}} - T_{\text{initial}}$ to determine the change in temperature.	$\begin{aligned}\Delta T &= T_{\text{final}} - T_{\text{initial}} \\ &= 45.00^\circ\text{C} - 10.00^\circ\text{C} \\ &= 35.00^\circ\text{C}\end{aligned}$
Use the formula $Q = mc\Delta T$ to determine the heat, $Q$ , gained by the hot chocolate and glass mug. Assume that the density of the hot chocolate is the same as the density of water.	$\begin{aligned}Q &= (mc\Delta T)_{\text{mug}} + (mc\Delta T)_{\text{chocolate}} \\ &= [(1.50 \times 10^2 \text{ g})(0.84 \text{ J/g}\cdot^\circ\text{C})(35.00^\circ\text{C})] \\ &\quad + [(5.00 \times 10^2 \text{ g})(3.75 \text{ J/g}\cdot^\circ\text{C})(35.00^\circ\text{C})] \\ &= 4410 \text{ J} + 65\,625 \text{ J} \\ &= 70\,035 \text{ J} \\ &= 70.035 \text{ kJ}\end{aligned}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find the enthalpy change, $\Delta H$ , during the combustion.	$\begin{aligned}\Delta H &= -Q \\ &= -70.035 \text{ kJ}\end{aligned}$
Use the formula $\Delta H_{\text{comb}} = \frac{\Delta H}{m}$ to calculate the enthalpy of combustion per gram of fuel.	$\begin{aligned}H_{\text{comb}} &= \frac{\Delta H}{m} \\ &= \frac{-70.035 \text{ kJ}}{2.75 \text{ g}} \\ &= -25.467 \text{ kJ/g}\end{aligned}$ <p>The enthalpy of combustion of the fuel is <math>-25 \text{ kJ/g}</math>.</p>

**Check Your Solution**

The negative sign in the answer indicates that energy is released during the burning of the fuel. This is reasonable. The answer has the correct number of significant digits.

36. Urea,  $\text{CO}(\text{NH}_2)_2(\text{s})$ , is widely used in various aspects of fertilizer manufacturing. A 4.28 g sample of urea is burned in a flame calorimeter. The calorimeter is made of aluminum and has a mass of 40.3 g. The 125.0 mL volume of water inside the calorimeter is at an initial temperature of  $5.5^\circ\text{C}$ . After the urea is completely combusted, the water temperature is  $96.0^\circ\text{C}$ . Use this information to calculate the enthalpy of combustion,  $\Delta H_{\text{comb}}$ , for urea. Refer to **Table 5.1** for the specific heat capacity of aluminum.

### What Is Required?

You need to calculate the enthalpy of combustion for urea.

### What Is Given?

You know the mass of urea:  $m_{\text{urea}} = 4.28 \text{ g}$

You know the mass of the aluminum calorimeter:  $m_{\text{calorimeter}} = 40.3 \text{ g}$

You know the specific heat capacity of  $\text{Al}(\text{s})$ :  $c_{\text{Al}} = 0.897 \text{ J/g}\cdot^\circ\text{C}$

You know the volume of water:  $V_{\text{H}_2\text{O}} = 125.0 \text{ mL}$

You know the mass of water:  $m_{\text{H}_2\text{O}} = 125.0 \text{ g}$

You know the specific heat capacity of  $\text{H}_2\text{O}(\ell)$ :  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^\circ\text{C}$

You know the initial temperature:  $T_{\text{initial}} = 5.5^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 96.0^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Determine the amount in moles, $n$ , of $\text{CO}(\text{NH}_2)_2(\text{s})$ combusted, using the mass and molar mass of urea and the formula $n = \frac{m}{M}$ .	$n = \frac{m}{M}$ $n_{\text{CH}_4\text{N}_2\text{O}} = \frac{4.28 \text{ g}}{60.07 \text{ g/mol}}$ $= 0.0712 \text{ mol}$
Use $\Delta T = T_{\text{final}} - T_{\text{initial}}$ to determine the change in temperature.	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 96.0^\circ\text{C} - 5.5^\circ\text{C}$ $= 90.5^\circ\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the amount of heat, $Q$ , absorbed by the flame calorimeter.	$Q = (mc\Delta T)_{\text{Al}} + (mc\Delta T)_{\text{H}_2\text{O}}$ $= [(40.3 \text{ g})(0.897 \text{ J/g}\cdot^\circ\text{C})(90.5^\circ\text{C})]$ $+ [(125.0 \times 10^3 \text{ g})(4.19 \text{ J/g}\cdot^\circ\text{C})(90.5^\circ\text{C})]$ $= 3271.5 \text{ J} + 47\,399.3 \text{ J}$ $= 50\,670.8 \text{ J}$ $= 50.671 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find the enthalpy change, $\Delta H$ , during the burning.	$\Delta H = -Q$ $= -50.671 \text{ kJ}$

<p>Use the formula <math>\Delta H_{\text{comb}}^{\circ} = \frac{\Delta H}{n}</math> to calculate the enthalpy of combustion.</p>	$\Delta H_{\text{comb}}^{\circ} = \frac{\Delta H}{n}$ $= \frac{-50.671 \text{ kJ}}{0.082 \text{ 07 mol}}$ $= -617.412 \text{ kJ/mol}$ <p>The enthalpy of combustion of urea is <math>-617 \text{ kJ/mol}</math>.</p>
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### Check Your Solution

The negative sign in the answer indicates that energy is released during the burning of the urea. This is reasonable. The answer has the correct number of significant digits.

37. When a sample of ethylene gas,  $\text{C}_2\text{H}_4(\text{g})$ , is burned in an open system, the heat produced warms 50.0 g of water by  $63.0^\circ\text{C}$ . The mass of  $\text{C}_2\text{H}_4(\text{g})$  burned was 0.268 g.
- What is the standard molar enthalpy of combustion for  $\text{C}_2\text{H}_4(\text{g})$  in  $\text{kJ/mol}$ ?
  - Write the thermochemical equation for this combustion reaction.

### What Is Required?

You need to calculate the standard molar enthalpy of combustion for ethylene gas,  $\text{C}_2\text{H}_4(\text{g})$  and write the thermochemical equation for the combustion reaction.

### What Is Given?

You know the mass of ethylene,  $\text{C}_2\text{H}_4(\text{g})$ :  $m_{\text{C}_2\text{H}_4} = 0.268 \text{ g}$

You know the mass of water:  $m_{\text{H}_2\text{O}} = 50.0 \text{ g}$

You know the change in temperature:  $\Delta T = 63.0^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
a. Determine the amount in moles, $n$ , of $\text{C}_2\text{H}_4(\text{g})$ burned, using the mass and molar mass of ethylene gas and the formula $n = \frac{m}{M}$ .	$n = \frac{m}{M}$ $n_{\text{C}_2\text{H}_4} = \frac{0.268 \cancel{\text{g}}}{28.06 \cancel{\text{g}}/\text{mol}}$ $= 0.009\,550\,9 \text{ mol}$
Use the formula $Q = mc\Delta T$ to calculate the amount of heat, $Q$ , absorbed by the water.	$Q = mc\Delta T$ $= (50.0 \cancel{\text{g}})(4.19 \text{ J}/\cancel{\text{g}}^\circ\text{C})(63^\circ\cancel{\text{C}})$ $= 13\,198.5 \text{ J}$ $= 13.198\,5 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find the enthalpy change, $\Delta H$ , during the burning.	$\Delta H = -Q$ $= -13.198\,5 \text{ kJ}$
Use the formula $\Delta H_{\text{comb}}^\circ = \frac{\Delta H}{n}$ to calculate the standard molar enthalpy of combustion.	$\Delta H_{\text{comb}}^\circ = \frac{\Delta H}{n}$ $= \frac{-13.198\,5 \text{ kJ}}{0.009\,550\,9 \text{ mol}}$ $= -1381.91 \text{ kJ/mol}$ <p>The standard molar enthalpy of combustion for <math>\text{C}_2\text{H}_4(\text{g})</math> is <math>-1.38 \times 10^3 \text{ kJ/mol}</math>.</p>
b. Write the thermochemical equation for the combustion reaction.	$\text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow$ $2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\ell) + 1.38 \times 10^3 \text{ kJ}$

### Check Your Solution

The negative sign in the answer indicates that energy is released during the burning of the ethylene. This is a reasonable. The answer has the correct number of significant digits.

38. Diborane,  $\text{B}_2\text{H}_6(\text{g})$ , is a highly reactive compound that has been used as a reducing agent in rocket propellants. It can react with chlorine gas as written below.



A 2.85 g sample of  $\text{B}_2\text{H}_6(\text{g})$  reacts with an excess of chlorine gas in a bomb calorimeter, which is calibrated to  $8.82 \text{ kJ}/^\circ\text{C}$ . What change in temperature will occur when the sample reacts?

### What Is Required?

You need to find the temperature change when a sample of diborane,  $\text{B}_2\text{H}_6(\text{g})$ , reacts.

### What Is Given?

You know the balanced thermochemical equation. From this, you know the enthalpy of reaction,  $\Delta H_r$ :  $-755 \text{ kJ}$

You know the mass,  $m$ , of  $\text{B}_2\text{H}_6(\text{g})$ :  $m_{\text{B}_2\text{H}_6} = 2.85 \text{ g}$

You know the heat capacity of the calorimeter:  $C = 8.82 \text{ kJ}/^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Determine the amount in moles, $n$ , of $\text{B}_2\text{H}_6(\text{g})$ that reacts, using the mass and molar mass of diborane and the formula $n = \frac{m}{M}$ .	$n = \frac{m}{M}$ $n_{\text{B}_2\text{H}_6} = \frac{2.85 \cancel{\text{g}}}{60.07 \cancel{\text{g}}/\text{mol}}$ $= 0.10296 \text{ mol}$
Use the equation, $\Delta H = n\Delta H_r$ to calculate the enthalpy change.	$\Delta H = n\Delta H_r$ $= (0.10296 \cancel{\text{mol}}) \left( -755 \frac{\text{kJ}}{\cancel{\text{mol}}} \right)$ $= -77.7348 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $\Delta H$ to find the amount of heat, $Q$ , absorbed by the calorimeter and its contents during the reaction.	$Q = -\Delta H$ $= 77.7348 \text{ kJ}$

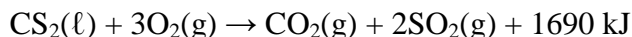


Use the formula $Q = C\Delta T$ to determine the change in temperature, $\Delta T$ .	$Q = C\Delta T$ $77.7348 \text{ kJ} = (8.82 \text{ kJ/}^\circ\text{C})\Delta T$ $\Delta T = \frac{77.7348 \cancel{\text{kJ}}}{8.82 \cancel{\text{kJ}}/^\circ\text{C}}$ $= 8.81^\circ\text{C}$ <p>The temperature is expected to increase by <math>8.81^\circ\text{C}</math>.</p>
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### Check Your Solution

The negative sign in the answer indicates that energy is released during the reaction of diborane and chlorine gas. This is expected since the temperature increased. This is a reasonable answer. The answer has the correct number of significant digits.

39. The equation below represents the burning of carbon disulfide,  $\text{CS}_2(\ell)$ , in a calorimeter. What mass of  $\text{CS}_2(\ell)$  must have burned to increase the temperature of 1.00 kg of water by  $15.5^\circ\text{C}$ ?



### What Is Required?

You need to find the mass of carbon disulfide,  $\text{CS}_2(\ell)$  that must have burned to cause a given change in temperature of water in a calorimeter.

### What Is Given?

You know the balanced thermochemical equation. From this, you know the enthalpy of combustion for  $\text{CS}_2(\ell)$ :  $\Delta H_{\text{comb}} = -1690 \text{ kJ/mol}$

You know the mass of water:  $m_{\text{H}_2\text{O}} = 1.00 \text{ kg}$

You know the change in temperature:  $\Delta T: 15.5^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Convert the mass of water, $m$ , from kilograms to grams.	$m = 1.00 \cancel{\text{ kg}} \times (1 \times 10^3 \text{ g} / \cancel{\text{ kg}})$ $= 1 \times 10^3 \text{ g}$
Use the equation $Q = mc\Delta T$ to calculate the amount of heat, $Q$ , absorbed by the water during the burning.	$Q = mc\Delta T$ $= (1.00 \times 10^3 \cancel{\text{ g}})(4.19 \text{ J} / \cancel{\text{ g}}^\circ\text{C})(15.5^\circ\text{C})$ $= 64\,945 \text{ J}$ $= 64.945 \text{ kJ}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find the enthalpy change, $\Delta H$ , during the burning.	$\Delta H = -Q$ $= -64.945 \text{ kJ}$
Use the formula $\Delta H_{\text{comb}} = \frac{\Delta H}{n}$ to determine the amount in moles, $n$ , of $\text{CS}_2(\ell)$ that burned.	$\Delta H_{\text{comb}} = \frac{\Delta H}{n}$ $-1690 \text{ kJ/mol} = \frac{-64.945 \text{ kJ}}{n_{\text{CS}_2}}$ $n_{\text{CS}_2} = \frac{-64.945 \cancel{\text{ kJ}}}{-1690 \cancel{\text{ kJ/mol}}}$ $= 0.038428 \text{ mol}$
Determine the mass, $m$ , of $\text{CS}_2(\ell)$ that burned, using the molar mass and amount in moles of carbon disulfide, and the formula $m = nM$ .	$m = nM$ $m_{\text{CS}_2} = (0.038428 \cancel{\text{ mol}})(76.15 \text{ g} / \cancel{\text{ mol}})$ $= 2.92629 \text{ g}$ <p>The mass of carbon disulfide that burned is 2.93 g.</p>

### Check Your Solution

The answer has the correct number of significant digits and seems to be a reasonable mass to produce enough thermal energy to heat 1.00 kg of water by  $15.5^\circ\text{C}$ .

40. A sample of acetone having a mass of 1.920 g undergoes complete combustion in an insulated calorimeter holding 400.0 mL of water. It is determined that the water and calorimeter together absorb 58.829 kJ of heat energy from this combustion. Based upon this information, what is the standard molar enthalpy of combustion for acetone in kJ/mol?

### What Is Required?

You need to calculate the standard molar enthalpy of combustion for acetone,  $\text{C}_3\text{H}_6\text{O}(\ell)$ .

### What Is Given?

You know the mass of acetone:  $m_{\text{acetone}} = 1.920 \text{ g}$

You know the volume of water:  $V_{\text{H}_2\text{O}} = 400.0 \text{ mL}$

You know the mass of water:  $m_{\text{H}_2\text{O}} = 400.0 \text{ g}$

You know the amount of heat gained by the water and calorimeter:  $Q = 58.829 \text{ kJ}$

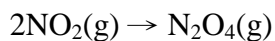
Plan Your Strategy	Act on Your Strategy
Determine the amount in moles, $n$ , of $\text{C}_3\text{H}_6\text{O}(\ell)$ combusted, using the mass and molar mass of acetone and the formula $n = \frac{m}{M}$ .	$n = \frac{m}{M}$ $n_{\text{C}_3\text{H}_6\text{O}} = \frac{1.920 \text{ g}}{58.09 \text{ g/mol}}$ $= 0.033 \text{ 05 mol}$
Since $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ , change the sign of $Q$ to find the enthalpy change, $\Delta H$ , during the combustion of the sample.	$\Delta H = -Q$ $= -58.829 \text{ kJ}$
Use the formula $\Delta H_{\text{comb}}^{\circ} = \frac{\Delta H}{n}$ to calculate the standard molar enthalpy of combustion.	$\Delta H_{\text{comb}}^{\circ} = \frac{\Delta H}{n}$ $= \frac{-58.829 \text{ kJ}}{0.033 \text{ 05 mol}}$ $= -1780 \text{ kJ/mol}$ <p>The standard enthalpy of combustion for acetone is <math>-1780 \text{ kJ/mol}</math>.</p>

### Check Your Solution

The negative sign in the answer indicates that energy is released during the burning of the acetone. This is a reasonable answer. The answer has the correct number of significant digits

## Using Hess's Law to Determine Enthalpy Change (Student textbook page 316)

41. Nitrogen dioxide,  $\text{NO}_2(\text{g})$ , is an emission resulting from the burning of gasoline in an automobile engine that contributes to the formation of smog and acid rain. It can be converted to dinitrogen tetroxide as shown below:



- a. Use Hess's law and the following equations to determine the enthalpy change for this reaction.



- b. Write the thermochemical equation for the overall reaction.

### What Is Required?

You need to determine the enthalpy change of a reaction by using two other reactions.

### What Is Given?

You are given the overall equation. You are given two equations and their corresponding enthalpy changes that can be manipulated to represent the individual steps that occur to reach the overall equation.

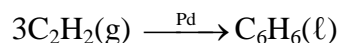
Plan Your Strategy	
<p>a. Apply Hess's law. The given equations can be manipulated so that their reactants and products match the reactants and products in the overall equation.</p> <p>To manipulate the equations, reverse (1).</p> <p>When an equation is reversed, the sign of <math>\Delta H^\circ</math> is reversed. If the stoichiometric coefficients in a given equation are multiplied by an integer or a fraction, multiply <math>\Delta H^\circ</math> for that equation by the same integer or fraction.</p>	
Act on Your Strategy	
$-1 \times (1): \quad 2\text{NO}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) \quad \Delta H^\circ = -66.4 \text{ kJ}$ $(2): \text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{N}_2\text{O}_4(\text{g}) \quad \Delta H^\circ = 11.1 \text{ kJ}$	
Plan Your Strategy	
<p>To determine the overall reaction, add the manipulated equations together, cancelling species that are common to both sides.</p> <p>Add the <math>\Delta H^\circ</math> values together to find the enthalpy change for the overall reaction.</p>	

Act on Your Strategy	
$  \begin{array}{rcl}  2\text{NO}_2(\text{g}) & \rightarrow & \cancel{\text{N}_2(\text{g})} + \cancel{2\text{O}_2(\text{g})} \\  \cancel{\text{N}_2(\text{g})} + \cancel{2\text{O}_2(\text{g})} & \rightarrow & \text{N}_2\text{O}_4(\text{g})  \end{array}  $	$  \begin{array}{rcl}  \Delta H^\circ & = & -66.4 \text{ kJ} \\  \Delta H^\circ & = & 11.1 \text{ kJ} \\  \hline  \Delta H^\circ & = & -55.3 \text{ kJ}  \end{array}  $
Plan Your Strategy	Act on Your Strategy
b. Write the thermochemical equation for the overall reaction.	$2\text{NO}_2(\text{g}) \rightarrow \text{N}_2\text{O}_4(\text{g}) + 55.3 \text{ kJ}$

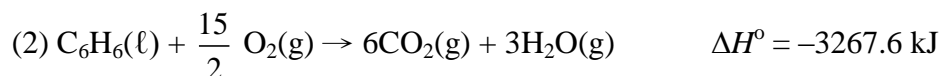
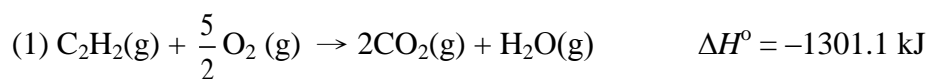
### Check Your Solution

Since the equations add to give the overall equation, the given individual equations must have been manipulated correctly. Check that the correct sign has been carried through with each  $\Delta H^\circ$  term. The overall  $\Delta H^\circ$  shows the same precision as the least precise used in the calculation with one digit to the right of the decimal point.

42. Ethyne,  $\text{C}_2\text{H}_2(\text{g})$ , can be converted to benzene,  $\text{C}_6\text{H}_6(\ell)$ , over a palladium catalyst.



Determine the enthalpy of reaction for this process from the equations below that show the combustion of  $\text{C}_2\text{H}_2(\text{g})$  and  $\text{C}_6\text{H}_6(\ell)$  at standard conditions.



### What Is Required?

You need to determine the enthalpy change of a reaction by using two other reactions.

### What Is Given?

You are given the overall equation. You are given two equations and their corresponding enthalpy changes that can be manipulated to represent the individual steps that occur to reach the overall equation.

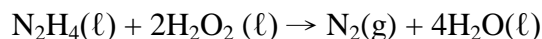
Plan Your Strategy	
Apply Hess's law. The given equations can be manipulated so that their reactants and products match the reactants and products in the overall equation.	
Manipulate the equations by multiplying the coefficients in (1) by 3 and reversing (2).	
When an equation is reversed, the sign of $\Delta H^\circ$ is reversed. If the stoichiometric coefficients in a given equation are multiplied by an integer or a fraction, multiply $\Delta H^\circ$ for that equation by the same integer or fraction.	
Act on Your Strategy	
$3 \times (1): 3\text{C}_2\text{H}_2(\text{g}) + \frac{15}{2} \text{O}_2(\text{g}) \rightarrow 6\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{g})$	$\Delta H^\circ = -3903.3 \text{ kJ}$
$-1 \times (2): 6\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{g}) \rightarrow \text{C}_6\text{H}_6(\ell)$	$\Delta H^\circ = 3267.6 \text{ kJ}$

Plan Your Strategy	
To determine the overall reaction, add the manipulated equations together, cancelling species that are common to both sides.	
Add the $\Delta H^\circ$ values together to find the enthalpy change for the overall reaction.	
Act on Your Strategy	
$3\text{C}_2\text{H}_2(\text{g}) + \cancel{\frac{15}{2}\text{O}_2(\text{g})} \rightarrow \cancel{6\text{CO}_2(\text{g})} + \cancel{3\text{H}_2\text{O}(\text{g})}$	$\Delta H^\circ = -3903.3 \text{ kJ}$
$\cancel{6\text{CO}_2(\text{g})} + \cancel{3\text{H}_2\text{O}(\text{g})} \rightarrow \text{C}_6\text{H}_6(\ell) + \cancel{\frac{15}{2}\text{O}_2(\text{g})}$	$\Delta H^\circ = 3267.6 \text{ kJ}$
<hr/> $\text{C}_2\text{H}_2(\text{g}) \rightarrow \text{C}_6\text{H}_6(\ell)$	
$\Delta H^\circ = -635.7 \text{ kJ}$	

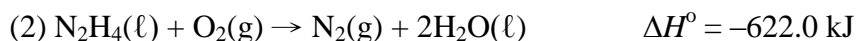
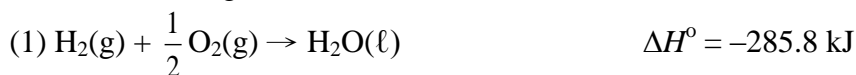
### Check Your Solution

Since the equations add to give the overall equation, the given individual equations must have been manipulated correctly. Check that the correct sign has been carried through with each  $\Delta H^\circ$  term. The overall  $\Delta H^\circ$  shows the same precision as the least precise used in the calculation with one digit to the right of the decimal point.

43. Hydrazine,  $\text{N}_2\text{H}_4(\ell)$ , is a high-energy compound used as a rocket propellant. Use Hess's law to determine the enthalpy of reaction when this compound reacts as follows:



Use the following information:



### What Is Required?

You need to determine the enthalpy change of a reaction by using three other reactions.

### What Is Given?

You are given the overall equation.

You are given three equations and their corresponding enthalpy changes that can be manipulated to represent the individual steps that occur to reach the overall equation.

Plan Your Strategy		
Apply Hess's law. The given equations can be manipulated so that their reactants and products match the reactants and products in the overall equation.		
Manipulate the equations by multiplying the coefficients in (1) by 2. Then, reverse (3) and multiply the coefficients in that equation by 2.		
When an equation is reversed, the sign of $\Delta H^\circ$ is reversed. If the stoichiometric coefficients in a given equation are multiplied by an integer or a fraction, multiply $\Delta H^\circ$ for that equation by the same integer or fraction.		
Act on Your Strategy		
$2 \times (1):$	$2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)$	$\Delta H^\circ = -571.6 \text{ kJ}$
$1 \times (2):$	$\text{N}_2\text{H}_4(\ell) + \text{O}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\ell)$	$\Delta H^\circ = -622.0 \text{ kJ}$
$-2 \times (3):$	$2\text{H}_2\text{O}_2(\ell) \rightarrow 2\text{H}_2(\text{g}) + 2\text{O}_2(\text{g})$	$\Delta H^\circ = 376.0 \text{ kJ}$

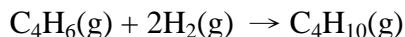


Plan Your Strategy	
To determine the overall reaction, add the manipulated equations together, cancelling species that are common to both sides.	
Add the $\Delta H^\circ$ values together to find the enthalpy change for the overall reaction.	
Act on Your Strategy	
$\cancel{2\text{H}_2(\text{g})} + \cancel{\text{O}_2(\text{g})} \rightarrow 2\text{H}_2\text{O}(\ell)$	$\Delta H^\circ = -571.6 \text{ kJ}$
$\text{N}_2\text{H}_4(\ell) + \cancel{\text{O}_2(\text{g})} \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\ell)$	$\Delta H^\circ = -622.0 \text{ kJ}$
$2\text{H}_2\text{O}_2(\ell) \rightarrow \cancel{2\text{H}_2(\text{g})} + \cancel{2\text{O}_2(\text{g})}$	$\Delta H^\circ = 376.0 \text{ kJ}$
$\text{N}_2\text{H}_4(\ell) + 2\text{H}_2\text{O}_2(\ell) \rightarrow \text{N}_2(\text{g}) + 4\text{H}_2\text{O}(\ell)$	$\Delta H^\circ = -817.6 \text{ kJ}$

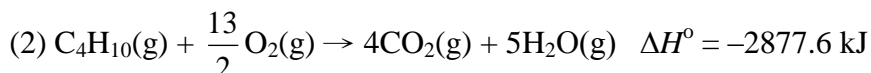
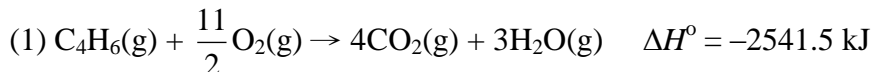
### Check Your Solution

Since the equations add to give the overall equation, the given individual equations must have been manipulated correctly. Check that the correct sign has been carried through with each  $\Delta H^\circ$  term. The overall  $\Delta H^\circ$  shows the same precision as the least precise used in the calculation with one digit to the right of the decimal point.

44. Synthetic rubber can be manufactured from the hydrocarbon 1,3-butadiene,  $\text{C}_4\text{H}_6(\text{g})$ . This compound reacts with hydrogen to produce butane(g),  $\text{C}_4\text{H}_{10}(\text{g})$ , as shown in the equation below:



Use Hess's law and the equations that follow to determine the enthalpy change for this reaction.



### What Is Required?

You need to determine the enthalpy change of a reaction by using three other reactions.

### What Is Given?

You are given the overall equation. You are given three equations and their corresponding enthalpy changes that can be manipulated to represent the individual steps that occur to reach the overall equation.

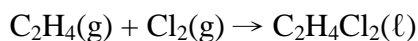
Plan Your Strategy	
Apply Hess's law. The given equations can be manipulated so that their reactants and products match the reactants and products in the overall equation.	
Manipulate the equations by reversing (2). Then, multiply the coefficients in (3) by 2.	
When an equation is reversed, the sign of $\Delta H^\circ$ is reversed. If the stoichiometric coefficients in a given equation are multiplied by an integer or a fraction, multiply $\Delta H^\circ$ for that equation by the same integer or fraction.	
Act on Your Strategy	
$1 \times (1): \text{C}_4\text{H}_6(\text{g}) + \frac{11}{2}\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{g})$	$\Delta H^\circ = -2541.5 \text{ kJ}$
$-1 \times (2): 4\text{CO}_2(\text{g}) + 5\text{H}_2\text{O}(\text{g}) \rightarrow \text{C}_4\text{H}_{10}(\text{g}) + \frac{13}{2}\text{O}_2(\text{g})$	$\Delta H^\circ = 2877.6 \text{ kJ}$
$2 \times (3): 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$	$\Delta H^\circ = -483.6 \text{ kJ}$
Plan Your Strategy	
To determine the overall reaction, add the manipulated equations together, cancelling species that are common to both sides.	
Add the $\Delta H^\circ$ values together to find the enthalpy change for the overall reaction.	

Act on Your Strategy		
$\text{C}_4\text{H}_6(\text{g}) + \frac{11}{2}\text{O}_2(\text{g}) \rightarrow \cancel{4\text{CO}_2(\text{g})} + \cancel{3\text{H}_2\text{O}(\text{g})}$	$\Delta H^\circ = -2541.5 \text{ kJ}$	
$\cancel{4\text{CO}_2(\text{g})} + \cancel{5\text{H}_2\text{O}(\text{g})} \rightarrow \text{C}_4\text{H}_{10}(\text{g}) + \frac{13}{2}\text{O}_2(\text{g})$	$\Delta H^\circ = 2877.6 \text{ kJ}$	
$2\text{H}_2(\text{g}) + \cancel{\text{O}_2(\text{g})} \rightarrow \cancel{2\text{H}_2\text{O}(\text{g})}$	$\Delta H^\circ = -483.6 \text{ kJ}$	
$\text{C}_4\text{H}_6(\text{g}) + 2\text{H}_2(\text{g}) \rightarrow \text{C}_4\text{H}_{10}(\text{g})$	$\Delta H^\circ = -147.5 \text{ kJ}$	

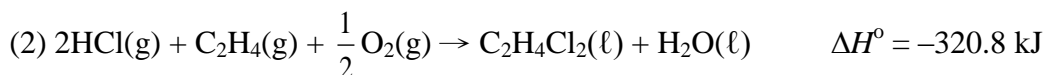
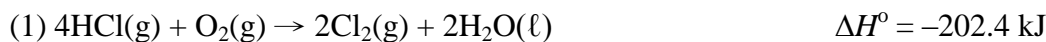
### Check Your Solution

Since the equations add to give the overall equation, the given individual equations must have been manipulated correctly. Check that the correct sign has been carried through with each  $\Delta H^\circ$  term. The overall  $\Delta H^\circ$  shows the same precision as the least precise used in the calculation with one digit to the right of the decimal point.

45. Ethylene,  $\text{C}_2\text{H}_4(\text{g})$ , is commonly used as an agent to hasten the ripening of fruit, such as bananas. It can also be used to prepare 1,2-dichloroethane,  $\text{C}_2\text{H}_4\text{Cl}_2(\ell)$ , which is used to make vinyl chloride. Use Hess's law to determine the enthalpy of reaction for the preparation of  $\text{C}_2\text{H}_4\text{Cl}_2(\ell)$ .



Given:



### What Is Required?

You need to determine the enthalpy change of a reaction by using two other reactions.

### What Is Given?

You are given the overall equation.

You are given two equations and their corresponding enthalpy changes that can be manipulated to represent the individual steps that occur to reach the overall equation.

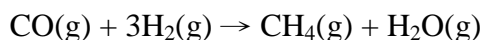
Plan Your Strategy		
Apply Hess's law. The given equations can be manipulated so that their reactants and products match the reactants and products in the overall equation.		
Manipulate the equations by reversing (1). Then, divide the coefficients of (1) by 2.		
When an equation is reversed, the sign of $\Delta H^\circ$ is reversed. If the stoichiometric coefficients in a given equation are multiplied by an integer or a fraction, multiply $\Delta H^\circ$ for that equation by the same integer or fraction.		
Act on Your Strategy		
$-\frac{1}{2} \times (1):$	$\text{Cl}_2(\text{g}) + \text{H}_2\text{O}(\ell) \rightarrow 2\text{HCl}(\text{g}) + \frac{1}{2}\text{O}_2(\text{g})$	$\Delta H^\circ = 101.2 \text{ kJ}$
(2):	$2\text{HCl}(\text{g}) + \text{C}_2\text{H}_4(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{C}_2\text{H}_4\text{Cl}_2(\ell) + \text{H}_2\text{O}(\ell)$	$\Delta H^\circ = -320.8 \text{ kJ}$
Plan Your Strategy		
To determine the overall reaction, add the manipulated equations together, cancelling species that are common to both sides.		
Add the $\Delta H^\circ$ values together to find the enthalpy change for the overall reaction.		

Act on Your Strategy	
$\text{Cl}_2(\text{g}) + \cancel{\text{H}_2\text{O}(\ell)} \rightarrow \cancel{2\text{HCl}(\text{g})} + \cancel{\frac{1}{2}\text{O}_2(\text{g})} \quad \Delta H^\circ = 101.2 \text{ kJ}$	
$\cancel{2\text{HCl}(\text{g})} + \text{C}_2\text{H}_4(\text{g}) + \cancel{\frac{1}{2}\text{O}_2(\text{g})} \rightarrow \text{C}_2\text{H}_4\text{Cl}_2(\ell) + \cancel{\text{H}_2\text{O}(\ell)} \quad \Delta H^\circ = -320.8 \text{ kJ}$	
<hr/> $\text{C}_2\text{H}_4(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{C}_2\text{H}_4\text{Cl}_2(\ell) \quad \Delta H^\circ = -219.6 \text{ kJ}$	

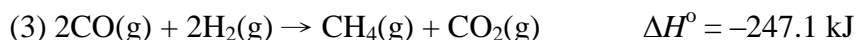
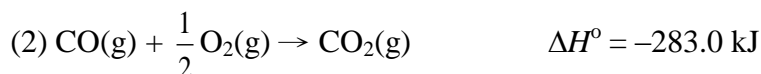
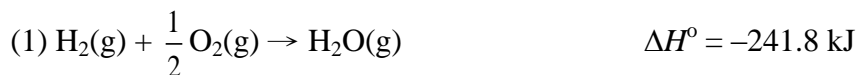
### Check Your Solution

Since the equations add to give the overall equation, the given individual equations must have been manipulated correctly. Check that the correct sign has been carried through with each  $\Delta H^\circ$  term. The overall  $\Delta H^\circ$  shows the same precision as the least precise used in the calculation with one digit to the right of the decimal point.

46. Carbon monoxide,  $\text{CO(g)}$ , can react with hydrogen,  $\text{H}_2\text{(g)}$ , to produce methane,  $\text{CH}_4\text{(g)}$ , and water vapour.



Given the following equations, use Hess's law to determine the enthalpy of reaction.



### What Is Required?

You need to determine the enthalpy change of a reaction by using three other reactions.

### What Is Given?

You are given the overall equation.

You are given three equations and their corresponding enthalpy changes that can be manipulated to represent the individual steps that occur to reach the overall equation.

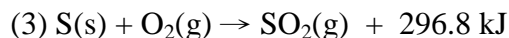
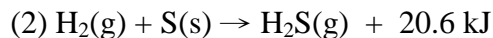
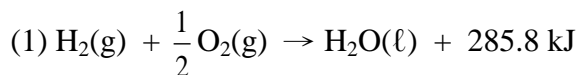
Plan Your Strategy	
Apply Hess's law. The given equations can be manipulated so that their reactants and products match the reactants and products in the overall equation.	
Manipulate the equations by reversing (2).	
When an equation is reversed, the sign of $\Delta H^\circ$ is reversed.	
Act on Your Strategy	
(1): $\text{H}_2\text{(g)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{H}_2\text{O(g)}$	$\Delta H^\circ = -241.8 \text{ kJ}$
$-1 \times (2): \text{CO}_2\text{(g)} \rightarrow \text{CO(g)} + \frac{1}{2}\text{O}_2\text{(g)}$	$\Delta H^\circ = 283.0 \text{ kJ}$
(3): $2\text{CO(g)} + 2\text{H}_2\text{(g)} \rightarrow \text{CH}_4\text{(g)} + \text{CO}_2\text{(g)}$	$\Delta H^\circ = -247.1 \text{ kJ}$
Plan Your Strategy	
To determine the overall reaction, add the manipulated equations together, cancelling species that are common to both sides.	
Add the $\Delta H^\circ$ values together to find the enthalpy change for the overall reaction.	

Act on Your Strategy	
$\text{H}_2(\text{g}) + \cancel{\frac{1}{2}\text{O}_2(\text{g})} \rightarrow \text{H}_2\text{O}(\text{g})$	$\Delta H^\circ = -241.8 \text{ kJ}$
$\cancel{\text{CO}_2(\text{g})} \rightarrow \cancel{\text{CO}(\text{g})} + \cancel{\frac{1}{2}\text{O}_2(\text{g})}$	$\Delta H^\circ = 283.0 \text{ kJ}$
$\cancel{2}\text{CO}(\text{g}) + 2\text{H}_2(\text{g}) \rightarrow \text{CH}_4(\text{g}) + \cancel{\text{CO}_2(\text{g})}$	$\Delta H^\circ = -247.1 \text{ kJ}$
<hr/>	
$\text{CO}(\text{g}) + 3\text{H}_2 \rightarrow \text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g})$	$\Delta H^\circ = -205.9 \text{ kJ}$

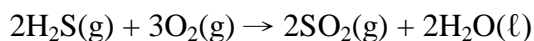
### Check Your Solution

Since the equations add to give the overall equation, the given individual equations must have been manipulated correctly. Check that the correct sign has been carried through with each  $\Delta H^\circ$  term. The overall  $\Delta H^\circ$  shows the same precision as the least precise used in the calculation with one digit to the right of the decimal point.

47. The following thermodynamic equations have been obtained from reference sources:



Use Hess's law to determine the enthalpy change for the reaction below:



### What Is Required?

You need to determine the enthalpy change of a reaction by using three other reactions.

### What Is Given?

You are given the overall equation.

You are given three equations and their corresponding enthalpy changes that can be manipulated to represent the individual steps that occur to reach the overall equation.

Plan Your Strategy	
Apply Hess's law. The given equations can be manipulated so that their reactants and products match the reactants and products in the overall equation.	
Manipulate the equations by first multiplying the coefficients of (1) by 2. Then, reverse (2) and multiply the coefficients of that equation by 2. Finally, multiply the coefficients of (3) by 2.	
When an equation is reversed, the sign of $\Delta H^\circ$ is reversed.	
Act on Your Strategy	
$2 \times (1): 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)$	$\Delta H^\circ = -571.6 \text{ kJ}$
$-2 \times (2): 2\text{H}_2\text{S}(\text{g}) \rightarrow 2\text{H}_2(\text{g}) + 2\text{S}(\text{g})$	$\Delta H^\circ = 41.2 \text{ kJ}$
$2 \times (3): 2\text{S}(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{SO}_2(\text{g})$	$\Delta H^\circ = -593.6 \text{ kJ}$

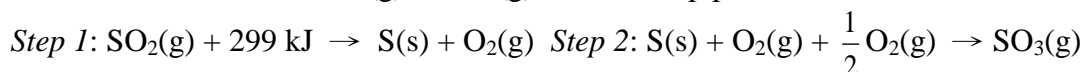


Plan Your Strategy	
To determine the overall reaction, add the manipulated equations together, cancelling species that are common to both sides.	
Add the $\Delta H^\circ$ values together to find the enthalpy change for the overall reaction.	
Act on Your Strategy	
<del>2H<sub>2</sub>(g)</del> + O <sub>2</sub> (g) → 2H <sub>2</sub> O(l)	$\Delta H^\circ = -571.6 \text{ kJ}$
2H <sub>2</sub> S(g) → <del>2H<sub>2</sub>(g)</del> + <del>2S(g)</del>	$\Delta H^\circ = 41.2 \text{ kJ}$
<del>2S(g)</del> + 2O <sub>2</sub> (g) → 2SO <sub>2</sub> (g)	$\Delta H^\circ = -593.6 \text{ kJ}$
2H <sub>2</sub> S(g) + 3O <sub>2</sub> (g) → 2SO <sub>2</sub> (g) + 2H <sub>2</sub> O(l)	$\Delta H^\circ = -1124.0 \text{ kJ}$

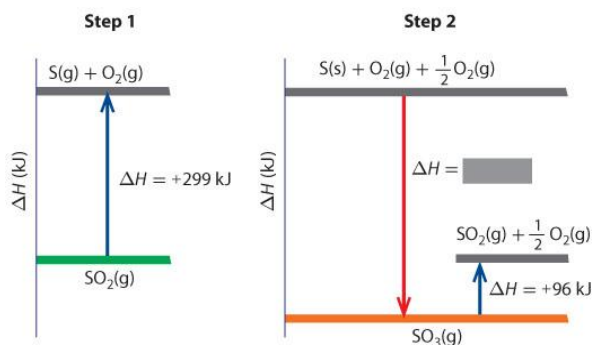
### Check Your Solution

Since the equations add to give the overall equation, the given individual equations must have been manipulated correctly. Check that the correct sign has been carried through with each  $\Delta H^\circ$  term. The overall  $\Delta H^\circ$  shows the same precision as the least precise used in the calculation with one digit to the right of the decimal point.

48. The reaction to convert  $\text{SO}_2(\text{g})$  to  $\text{SO}_3(\text{g})$  is a two-step process:



The enthalpy diagram below is a graphical representation of the process. Use this diagram to determine the enthalpy change for step 2.



### What Is Required?

You need to determine the enthalpy change for one step in a reaction sequence.

### What Is Given?

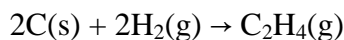
You have a graphical representation of the two steps in the sequence and the overall reaction.

Plan Your Strategy	
Apply Hess's law. The overall reaction is determined by adding the equations in steps 1 and 2 together, with species that are common to both sides being cancelled.	
Act on Your Strategy	
(1): $\text{SO}_2(\text{g}) \rightarrow \cancel{\text{S}(\text{s})} + \cancel{\text{O}_2(\text{g})}$	$\Delta H_1^\circ = +299 \text{ kJ}$
(2): $\cancel{\text{S}(\text{s})} + \cancel{\text{O}_2(\text{g})} + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{SO}_3(\text{g})$	$\Delta H_2^\circ = x \text{ kJ}$
$\text{SO}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{SO}_3(\text{g})$	$\Delta H_3^\circ = -96.0 \text{ kJ}$
Plan Your Strategy	Act on Your Strategy
The enthalpy change for the overall reaction is the sum of the $\Delta H^\circ$ values of each step. Set up an equation and solve for $x$ to determine the enthalpy change for step 2.	$\Delta H_1^\circ + \Delta H_2^\circ = \Delta H_3^\circ$ $(+299 \text{ kJ}) + (x \text{ kJ}) = (-96 \text{ kJ})$ $(x \text{ kJ}) = (-96 \text{ kJ}) - (+299 \text{ kJ})$ $x = -397$ <p>The enthalpy change for step 2 is <math>-397 \text{ kJ}</math>.</p>

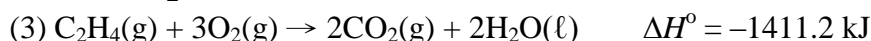
### Check Your Solution

The algebraic solution,  $\Delta H_2 = \Delta H_3 - \Delta H_1$ , corresponds to the mathematical depiction of the reaction.

49. Ethene,  $\text{C}_2\text{H}_4(\text{g})$ , is used in the manufacture of many polymers. If ethene could be formed from the elements carbon and hydrogen, the equation would be as follows:



Use Hess's law and the equations given below to determine the molar enthalpy of formation for  $\text{C}_2\text{H}_4(\text{g})$ .



### What Is Required?

You need to determine the enthalpy change of a reaction by using three other reactions.

### What Is Given?

You are given the overall equation.

You are given three equations and their corresponding enthalpy changes that can be manipulated to represent the individual steps that occur to reach the overall equation.

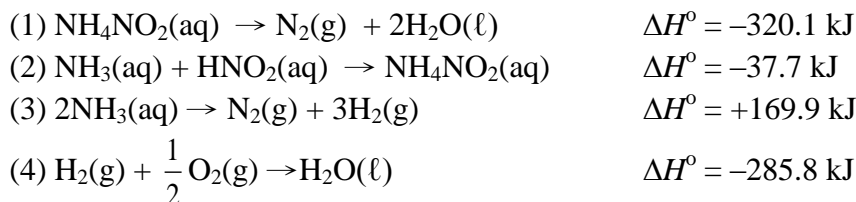
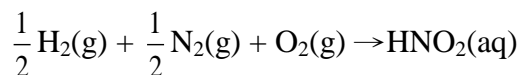
Plan Your Strategy		
Apply Hess's law. The given equations can be manipulated so that their reactants and products match the reactants and products in the overall equation.		
Manipulate the equations by first multiplying the coefficients of each of (1) and (2) by 2. The, reverse (3).		
When an equation is reversed, the sign of $\Delta H^\circ$ is reversed.		
Act on Your Strategy		
$2 \times (1):$	$2\text{C}(\text{s}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g})$	$\Delta H^\circ = -787.0 \text{ kJ}$
$2 \times (2):$	$2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)$	$\Delta H^\circ = -571.6 \text{ kJ}$
$-1 \times (3):$	$2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\ell) \rightarrow \text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g})$	$\Delta H^\circ = 1411.2 \text{ kJ}$

Plan Your Strategy	
To determine the overall reaction, add the manipulated equations together, cancelling species that are common to both sides.	
Add the $\Delta H^\circ$ values together to find the enthalpy change for the overall reaction.	
Act on Your Strategy	
$2\text{C(s)} + \cancel{2\text{O}_2\text{(g)}} \rightarrow \cancel{2\text{CO}_2\text{(g)}}$	$\Delta H^\circ = -787.0 \text{ kJ}$
$2\text{H}_2\text{(g)} + \cancel{\text{O}_2\text{(g)}} \rightarrow \cancel{2\text{H}_2\text{O(l)}}$	$\Delta H^\circ = -571.6 \text{ kJ}$
$-1 \times (3): \quad 2\text{CO}_2\text{(g)} + 2\text{H}_2\text{O(l)} \rightarrow \text{C}_2\text{H}_4\text{(g)} + 3\text{O}_2\text{(g)}$	$\Delta H^\circ = 1411.2 \text{ kJ}$
$2\text{C(s)} + 2\text{H}_2\text{(g)} \rightarrow \text{C}_2\text{H}_2\text{(g)}$	$\Delta H^\circ_f = 52.6 \text{ kJ}$

### Check Your Solution

Since the equations add to give the overall equation, the given individual equations must have been manipulated correctly. Check that the correct sign has been carried through with each  $\Delta H^\circ$  term. The overall  $\Delta H^\circ$  shows the same precision as the least precise used in the calculation with one digit to the right of the decimal point.

50. From the following equations, determine the molar enthalpy of formation for  $\text{HNO}_2(\text{aq})$ , as shown below in the overall equation:



### What Is Required?

You need to determine the enthalpy change of a reaction by using four other reactions.

### What Is Given?

You are given the overall equation.

You are given four equations and their corresponding enthalpy changes that can be manipulated to represent the individual steps that occur to reach the overall equation.

Plan Your Strategy	
Apply Hess's law. The given equations can be manipulated so that their reactants and products match the reactants and products in the overall equation.	
Manipulate the equations by first reversing each of (1) and (2). Then, multiply the coefficients of (3) by $\frac{1}{2}$ and multiply the coefficients of (4) by 2.	
When an equation is reversed, the sign of $\Delta H^\circ$ is reversed.	
Act on Your Strategy	
$-1 \times (1): \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\ell) \rightarrow \text{NH}_4\text{NO}_2(\text{aq})$	$\Delta H^\circ = +320.1 \text{ kJ}$
$-1 \times (2): \text{NH}_4\text{NO}_2(\text{aq}) \rightarrow \text{NH}_3(\text{aq}) + \text{HNO}_2(\text{aq})$	$\Delta H^\circ = +37.7 \text{ kJ}$
$\frac{1}{2} \times (3): \text{NH}_3(\text{aq}) \rightarrow \frac{1}{2} \text{N}_2(\text{g}) + \frac{3}{2} \text{H}_2(\text{g})$	$\Delta H^\circ = +84.95 \text{ kJ}$
$2 \times (4): 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)$	$\Delta H^\circ = -571.6 \text{ kJ}$

Plan Your Strategy	
To determine the overall reaction, add the manipulated equations together, cancelling species that are common to both sides.	
Add the $\Delta H^\circ$ values together to find the enthalpy change for the overall reaction.	
Act on Your Strategy	
<del><math>\text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\ell) \rightarrow \text{NH}_4\text{NO}_2(\text{aq})</math></del>	$\Delta H^\circ = +320.1 \text{ kJ}$
<del><math>\text{NH}_4\text{NO}_2(\text{aq}) \rightarrow \text{NH}_3(\text{aq}) + \text{HNO}_2(\text{aq})</math></del>	$\Delta H^\circ = +37.7 \text{ kJ}$
<del><math>\text{NH}_3(\text{aq}) \rightarrow \frac{1}{2}\text{N}_2(\text{g}) + \frac{3}{2}\text{H}_2(\text{g})</math></del>	$\Delta H^\circ = +84.95 \text{ kJ}$
<del><math>2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)</math></del>	$\Delta H^\circ = -571.6 \text{ kJ}$
$\frac{1}{2}\text{H}_2(\text{g}) + \frac{1}{2}\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{HNO}_2(\text{g})$	$\Delta H^\circ_f = -128.8 \text{ kJ}$

### Check Your Solution

Since the equations add to give the overall equation, the given individual equations must have been manipulated correctly. Check that the correct sign has been carried through with each  $\Delta H^\circ$  term. The overall  $\Delta H^\circ$  shows the same precision as the least precise used in the calculation with one digit to the right of the decimal point.

**Using Enthalpies of Formation**  
**Using an Enthalpy of Combustion to Determine an**  
**Enthalpy of Formation**  
**(Student textbook page 323)**

51. The complete combustion of ethanol,  $\text{C}_2\text{H}_5\text{OH}(\ell)$ , yields carbon dioxide and water vapour.
- Write the balanced chemical equation for this reaction.
  - Use enthalpies of formation to determine the enthalpy of combustion,  $\Delta H_{\text{comb}}^\circ$ , in kilojoules per mole of  $\text{C}_2\text{H}_5\text{OH}(\ell)$ .

**What Is Required?**

You need to calculate the enthalpy of combustion for the reaction.

**What Is Given?**

From Appendix B (student textbook page 743), you can obtain the molar enthalpies of formation.

$$\Delta H_f^\circ \text{ of } \text{C}_2\text{H}_5\text{OH}(\ell) = -277.6 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$$

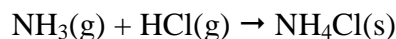
$$\Delta H_f^\circ \text{ of } \text{H}_2\text{O}(\text{g}) = -241.8 \text{ kJ/mol}$$

Plan Your Strategy	
a. Write the balanced chemical equation for the reaction.	
b. Use the available $\Delta H_f^\circ$ data to substitute into the equation $\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ , multiplying by the molar coefficients.	
Act on Your Strategy	
a) $\text{C}_2\text{H}_5\text{OH}(\ell) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{g})$	
b)	$\Delta H_{\text{comb}}^\circ \text{C}_2\text{H}_5\text{OH}(\ell) = \left[ (2 \text{ mol})(\Delta H_f^\circ \text{CO}_2(\text{g})) + (3 \text{ mol})(\Delta H_f^\circ \text{H}_2\text{O}(\text{g})) \right]$ $- \left[ (1 \text{ mol})(\Delta H_f^\circ \text{C}_2\text{H}_5\text{OH}(\ell)) + (3 \text{ mol})(\Delta H_f^\circ \text{O}_2(\text{g})) \right]$ $= \left[ (2 \text{ mol})(-393.5 \text{ kJ/mol}) + (3 \text{ mol})(-241.8 \text{ kJ/mol}) \right]$ $- \left[ (1 \text{ mol})(-277.6 \text{ kJ/mol}) + (3 \text{ mol})(0 \text{ kJ/mol}) \right]$ $= -787.0 \text{ kJ} - 725.4 \text{ kJ} + 277.6 \text{ kJ}$ $= -1234.8 \text{ kJ}$
The enthalpy of combustion of ethanol, $\text{C}_2\text{H}_5\text{OH}(\ell)$ , is $-1235 \text{ kJ/mol}$ .	

**Check Your Solution**

Check that substitution of the data has been carried out correctly. The answer shows the same precision as the given data, one digit to the right of the decimal.

52. A reagent bottle containing concentrated aqueous ammonia,  $\text{NH}_3(\text{aq})$ , is accidentally left open near a bottle of concentrated hydrochloric acid,  $\text{HCl}(\text{aq})$ . The cloud of white smoke that forms in the air near the two containers is ammonium chloride,  $\text{NH}_4\text{Cl}(\text{s})$ . The following reaction occurs in the air:



Use the standard molar heats of formation to calculate the enthalpy of reaction.

### What Is Required?

You need to calculate the enthalpy of reaction.

### What Is Given?

You know the balanced chemical equation for the reaction.

From Appendix B (student textbook page 743), you can obtain the molar enthalpies of formation.

$$\Delta H_f^\circ \text{ of } \text{NH}_3(\text{g}) = -45.9 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } \text{HCl}(\text{g}) = -92.3 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } \text{NH}_4\text{Cl}(\text{s}) = -314.4 \text{ kJ/mol}$$

Plan Your Strategy
Use the available $\Delta H_f^\circ$ data to substitute into the equation $\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ , multiplying by the molar coefficients.
Act on Your Strategy
$\begin{aligned} \Delta H_r^\circ &= \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants}) \\ &= [(1 \text{ mol})(\Delta H_f^\circ \text{NH}_4\text{Cl}(\text{s}))] - [(1 \text{ mol})(\Delta H_f^\circ \text{NH}_3(\text{g})) + (1 \text{ mol})(\Delta H_f^\circ \text{HCl}(\text{g}))] \\ &= [(1 \cancel{\text{ mol}})(-314.4 \text{ kJ}/\cancel{\text{ mol}})] \\ &\quad - [(1 \text{ mol})(-45.9 \text{ kJ}/\cancel{\text{ mol}}) + (1 \cancel{\text{ mol}})(-92.3 \text{ kJ}/\cancel{\text{ mol}})] \\ &= -314.4 \text{ kJ} - (-138.2 \text{ kJ}) \\ &= -314.4 \text{ kJ} + 138.2 \text{ kJ} \\ &= -176.2 \text{ kJ} \end{aligned}$ <p>The enthalpy of reaction is <math>-176.2 \text{ kJ}</math>.</p>

### Check Your Solution

Check that substitution of the data has been carried out correctly. The answer shows the same precision as the given data, one digit to the right of the decimal.

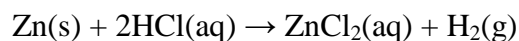


**53.** A piece of zinc reacts completely with hydrochloric acid,  $\text{HCl(aq)}$ , to produce an aqueous solution of zinc chloride,  $\text{ZnCl}_2\text{(aq)}$ , and hydrogen gas.

**a.** Use the data shown below to determine the enthalpy of reaction per mole of zinc.

$$\Delta H_f^\circ \text{ of } \text{HCl(aq)} = -167.2 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } \text{ZnCl}_2\text{(aq)} = -488.2 \text{ kJ/mol}$$



**b.** What mass of zinc must be used for 123 kJ of heat energy to be given off?

### What Is Required?

**a.** You need to determine the enthalpy of reaction per mole of  $\text{Zn(s)}$ .

**b.** You need to calculate the mass of zinc that will produce a given amount of heat.

### What Is Given?

**a.** You know the balanced chemical equation for the reaction.

You know the molar enthalpies of formation of  $\text{HCl(aq)}$  and  $\text{ZnCl}_2\text{(aq)}$ .

**b.** You know the mole ratio in the balanced chemical equation.

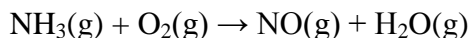
Plan Your Strategy
<b>a.</b> Use the available $\Delta H_f^\circ$ data to substitute into the equation $\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ , multiplying by the molar coefficients.
Act on Your Strategy
$\begin{aligned} \Delta H_r^\circ &= \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants}) \\ &= \left[ (1 \text{ mol})(\Delta H_f^\circ \text{ZnCl}_2\text{(aq)}) + (1 \text{ mol})(\Delta H_f^\circ \text{H}_2\text{(g)}) \right] \\ &\quad - \left[ (1 \text{ mol})(\Delta H_f^\circ \text{Zn(s)}) + (2 \text{ mol})(\Delta H_f^\circ \text{HCl(aq)}) \right] \\ &= \left[ (1 \text{ mol})(-488.2 \text{ kJ/mol}) + (1 \text{ mol})(0 \text{ kJ/mol}) \right] \\ &\quad - \left[ (1 \text{ mol})(0 \text{ kJ/mol}) + (2 \text{ mol})(-167.2 \text{ kJ/mol}) \right] \\ &= -488.2 \text{ kJ} - (-334.4 \text{ kJ}) \\ &= -488.2 \text{ kJ} + 334.4 \text{ kJ} \\ &= -153.8 \text{ kJ} \end{aligned}$ <p>The enthalpy of reaction is <math>-153.8 \text{ kJ}</math>.</p>

Plan Your Strategy	Act on Your Strategy
<b>b.</b> Use the equation $\Delta H = n\Delta H_r$ to calculate the amount in moles, $n$ , of Zn.	$\Delta H = n\Delta H_r$ $n = \frac{\Delta H}{\Delta H_r}$ $= \frac{-123 \text{ kJ}}{-153.8 \text{ kJ}}$ $= 0.799\,974 \text{ mol}$
Using the amount in moles and molar mass of Zn, use the formula $m = nM$ to determine the mass, $m$ , of zinc.	$m = nM$ $m_{\text{Zn}} = (0.799\,974 \text{ mol})(65.38 \text{ g/mol})$ $= 52.302 \text{ g}$ <p>The zinc would have a mass of 52.3 g.</p>

### Check Your Solution

- a.** Check that substitution of the data has been carried out correctly. The answer shows the same precision as the given data, one digit to the right of the decimal.
- b.** Less than a mol of zinc was needed. The answer seems reasonable since the amount of heat produced was less than the heat of reaction per mol of zinc. The answer has the correct number of significant digits.

54. Ammonia reacts with oxygen according to the following unbalanced equation:



Use standard molar enthalpies of formation to find the enthalpy of reaction for this reaction.

### What Is Required?

You need to find the enthalpy of reaction.

### What Is Given?

You know the unbalanced chemical equation.

From Appendix B (student textbook page 743), you can obtain the molar enthalpies of formation.

$$\Delta H_f^\circ \text{ of } \text{NH}_3(\text{g}) = -45.9 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } \text{NO}(\text{g}) = +91.3 \text{ kJ/mol}$$

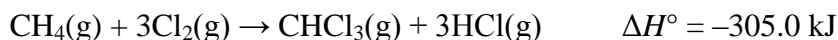
$$\Delta H_f^\circ \text{ of } \text{H}_2\text{O}(\text{g}) = -241.8 \text{ kJ/mol}$$

Plan Your Strategy	Act on Your Strategy
Write and balance the chemical equation for the reaction of ammonia with oxygen.	$4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g})$
Plan Your Strategy	
Use the available $\Delta H_f^\circ$ data to substitute into the equation $\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ , multiplying by the molar coefficients.	
Act on Your Strategy	
$\begin{aligned} \Delta H_r^\circ &= \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants}) \\ &= \left[ (4 \text{ mol})(\Delta H_f^\circ \text{NO}(\text{g})) + (6 \text{ mol})(\Delta H_f^\circ \text{H}_2\text{O}(\text{g})) \right] \\ &\quad - \left[ (4 \text{ mol})(\Delta H_f^\circ \text{NH}_3(\text{g})) + (5 \text{ mol})(\Delta H_f^\circ \text{O}_2(\text{g})) \right] \\ &= \left[ (4 \cancel{\text{mol}})(+91.3 \text{ kJ}/\cancel{\text{mol}}) + (6 \cancel{\text{mol}})(-241.8 \text{ kJ}/\cancel{\text{mol}}) \right] \\ &\quad - \left[ (4 \text{ mol})(-45.9 \text{ kJ}/\cancel{\text{mol}}) + (5 \cancel{\text{mol}})(0 \text{ kJ}/\cancel{\text{mol}}) \right] \\ &= (+365.2 \text{ kJ} - 1450.8 \text{ kJ}) - (-183.6 \text{ kJ}) \\ &= -1085.6 \text{ kJ} + 138.2 \text{ kJ} \\ &= -902 \text{ kJ} \end{aligned}$ <p>The enthalpy of reaction is <math>-902 \text{ kJ}</math></p>	

### Check Your Solution

Check that substitution of the data has been carried out correctly. The answer shows the same precision as the given data, one digit to the right of the decimal.

55. Consider the following equation representing the reaction of methane and chlorine to form chloroform,  $\text{CHCl}_3(\text{g})$ :



Use standard molar enthalpies of formation to determine the molar enthalpy of formation for chloroform.

### What Is Required?

You need to calculate the molar enthalpy of formation for chloroform.

### What Is Given?

You know the balanced chemical equation and the standard heat of reaction for the reaction.

From Appendix B (student textbook page 743), you can obtain the molar enthalpies of formation.

$$\Delta H_f^\circ \text{ of } \text{HCl}(\text{g}) = -92.3 \text{ kJ/mol}$$

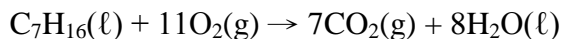
$$\Delta H_f^\circ \text{ of } \text{CH}_4(\text{g}) = -74.6 \text{ kJ/mol}$$

Plan Your Strategy
Use the available $\Delta H_f^\circ$ data to substitute into the equation $\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ , multiplying by the molar coefficients. Then solve for $\Delta H_f^\circ$ of $\text{CHCl}_3(\text{g})$ .
Act on Your Strategy
$\begin{aligned} \Delta H_r^\circ &= \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants}) \\ &= [(1 \text{ mol})(\Delta H_f^\circ \text{CHCl}_3(\text{g})) + (3 \text{ mol})(\Delta H_f^\circ \text{HCl}(\text{g}))] \\ &\quad - [(1 \text{ mol})(\Delta H_f^\circ \text{CH}_4(\text{g})) + (3 \text{ mol})(\Delta H_f^\circ \text{Cl}_2(\text{g}))] \\ -305.0 \text{ kJ} &= [(1 \text{ mol})(\Delta H_f^\circ \text{CHCl}_3(\text{g})) + (3 \text{ mol})(-92.3 \text{ kJ/mol})] \\ &\quad - [(1 \text{ mol})(-74.6 \text{ kJ/mol}) + (3 \text{ mol})(0 \text{ kJ/mol})] \\ &= (1 \text{ mol})(\Delta H_f^\circ \text{CHCl}_3(\text{g})) - 276.9 \text{ kJ} + 74.6 \text{ kJ} \\ &= (1 \text{ mol})(\Delta H_f^\circ \text{CHCl}_3(\text{g})) - 202.3 \text{ kJ} \\ \Delta H_f^\circ \text{CHCl}_3(\text{g}) &= -305.0 \text{ kJ} + 202.3 \text{ kJ} \\ \Delta H_f^\circ \text{CHCl}_3(\text{g}) &= -102.7 \text{ kJ/mol} \end{aligned}$ <p>The molar enthalpy of formation of chloroform is <math>-102.7 \text{ kJ/mol}</math>.</p>

### Check Your Solution

Check to see that substitution of the data has been carried out correctly. The answer shows the same precision as the given data, one digit to the right of the decimal.

56. The molar enthalpy of combustion for heptane,  $\text{C}_7\text{H}_{16}(\ell)$ , in a bomb calorimeter is  $-4816.7 \text{ kJ/mol}$  of heptane. Using this and  $\Delta H_f^\circ$  data, determine the molar enthalpy of formation of heptane.



### What Is Required?

You need to find the molar enthalpy of formation for heptane.

### What Is Given?

You know the balanced equation and the standard heat of reaction for the reaction. From Appendix B (student textbook page 743), you can obtain the molar enthalpies of formation.

$$\Delta H_f^\circ \text{ of } \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$$

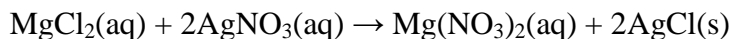
$$\Delta H_f^\circ \text{ of } \text{H}_2\text{O}(\ell) = -285.8 \text{ kJ/mol}$$

Plan Your Strategy
Use the available $\Delta H_f^\circ$ data and $\Delta H_{\text{comb}}^\circ$ to substitute into the equation $\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ , multiplying by the molar coefficients. Since the overall reaction is a combustion reaction, use $\Delta H_{\text{comb}}^\circ$ for $\Delta H_r^\circ$ .
Act on Your Strategy
$\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ $\Delta H_{\text{comb}}^\circ \text{C}_7\text{H}_{16}(\ell) = \left[ (7 \text{ mol})(\Delta H_f^\circ \text{CO}_2(\text{g})) + (8 \text{ mol})(\Delta H_f^\circ \text{H}_2\text{O}(\ell)) \right]$ $- \left[ (1 \text{ mol})(\Delta H_f^\circ \text{C}_7\text{H}_{16}(\ell)) + (11 \text{ mol})(\Delta H_f^\circ \text{O}_2(\text{g})) \right]$ $-4816.7 \text{ kJ} = \left[ (7 \text{ mol})(-393.5 \text{ kJ/mol}) + (8 \text{ mol})(-285.8 \text{ kJ/mol}) \right]$ $- \left[ (1 \text{ mol})(\Delta H_f^\circ \text{C}_7\text{H}_{16}(\ell)) + (11 \text{ mol})(0 \text{ kJ/mol}) \right]$ $= -2754.5 \text{ kJ} - 2286.4 \text{ kJ} - (1 \text{ mol})(\Delta H_f^\circ \text{C}_7\text{H}_{16}(\ell))$ $(1 \text{ mol})(\Delta H_f^\circ \text{C}_7\text{H}_{16}(\ell)) = -2754.5 \text{ kJ} - 2286.4 \text{ kJ} + 4816.7 \text{ kJ}$ $\Delta H_f^\circ \text{C}_7\text{H}_{16}(\ell) = -224.3 \text{ kJ/mol}$ <p>The molar enthalpy of formation for heptane is <math>-224.2 \text{ kJ/mol}</math>.</p>

### Check Your Solution

Check to see that substitution of the data has been carried out correctly. The answer shows the same precision as the given data, one digit to the right of the decimal.

57. A double displacement reaction occurs when solutions of magnesium chloride,  $\text{MgCl}_2(\text{aq})$ , and silver nitrate,  $\text{AgNO}_3(\text{aq})$ , are mixed. Assume that the reaction goes to completion and that all of the silver chloride that forms precipitates as  $\text{AgCl}(\text{s})$ .



Use the given enthalpy of formation data to determine the enthalpy of reaction.

$$\Delta H_f^\circ \text{ of } \text{MgCl}_2(\text{aq}) = -801.2 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } \text{Mg}(\text{NO}_3)_2(\text{aq}) = -881.6 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } \text{AgCl}(\text{s}) = -127.0 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } \text{AgNO}_3(\text{aq}) = -101.8 \text{ kJ/mol}$$

### What Is Required?

You need to find the enthalpy of reaction for a double displacement reaction.

### What Is Given?

You know the balanced chemical equation.

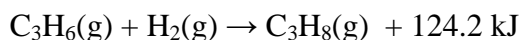
You are given the molar enthalpies of formation.

Plan Your Strategy
Use the available $\Delta H_f^\circ$ data to substitute into the equation $\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ , multiplying by the molar coefficients.
Act on Your Strategy
$\begin{aligned} \Delta H_r^\circ &= \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants}) \\ &= [(1 \text{ mol})(\Delta H_f^\circ \text{Mg}(\text{NO}_3)_2(\text{aq})) + (2 \text{ mol})(\Delta H_f^\circ \text{AgCl}(\text{s}))] \\ &\quad - [(1 \text{ mol})(\Delta H_f^\circ \text{MgCl}_2(\text{aq})) + (2 \text{ mol})(\Delta H_f^\circ \text{AgNO}_3(\text{aq}))] \\ &= [(1 \cancel{\text{mol}})(-881.6 \text{ kJ}/\cancel{\text{mol}}) + (2 \cancel{\text{mol}})(-127.0 \text{ kJ}/\cancel{\text{mol}})] \\ &\quad - [(1 \cancel{\text{mol}})(-801.2 \text{ kJ}/\cancel{\text{mol}}) + (2 \cancel{\text{mol}})(-101.8 \text{ kJ}/\cancel{\text{mol}})] \\ &= (-881.6 \text{ kJ} - 254.0 \text{ kJ}) - (-801.2 \text{ kJ} - 203.6 \text{ kJ}) \\ &= -1135.6 \text{ kJ} + 1018.8 \text{ kJ} \\ &= -130.8 \text{ kJ} \end{aligned}$ <p>The enthalpy of reaction is <math>-130.8 \text{ kJ}</math>.</p>

### Check Your Solution

Check that substitution of the data has been carried out correctly. The answer shows the same precision as the given data, one digit to the right of the decimal.

58. Hydrogen gas reacts with propene gas,  $\text{C}_3\text{H}_6(\text{g})$ , in an addition reaction to produce propane,  $\text{C}_3\text{H}_8(\text{g})$ .



Use enthalpy of formation data and the information given in the above equation to determine the enthalpy of formation of propene.

### What Is Required?

You need to find the enthalpy of formation for propene.

### What Is Given?

You know the balanced equation and the standard enthalpy of reaction for the reaction:  $\Delta H_r^\circ = -124.2 \text{ kJ}$

From Appendix B (student textbook page 743), you can obtain the molar enthalpy of formation of propane.

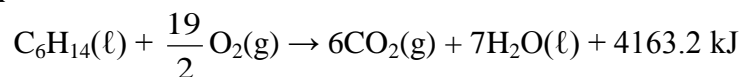
$$\Delta H_f^\circ \text{ of } \text{C}_3\text{H}_8(\text{g}) = -103.8 \text{ kJ/mol}$$

Plan Your Strategy
Use the available $\Delta H_f^\circ$ data to substitute into the equation $\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ , multiplying by the molar coefficients. Then solve for $\Delta H_f^\circ$ of $\text{C}_3\text{H}_6(\text{g})$ .
Act on Your Strategy
$\begin{aligned} \Delta H_r^\circ &= \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants}) \\ &= [(1 \text{ mol})(\Delta H_f^\circ \text{C}_3\text{H}_8(\text{g}))] \\ &\quad - [(1 \text{ mol})(\Delta H_f^\circ \text{C}_3\text{H}_6(\text{g})) + (1 \text{ mol})(\Delta H_f^\circ \text{H}_2(\text{g}))] \\ -124.2 \text{ kJ} &= [(1 \text{ mol})(-103.8 \text{ kJ/mol})] \\ &\quad - [(1 \text{ mol})(\Delta H_f^\circ \text{C}_3\text{H}_6(\text{g})) + (1 \text{ mol})(0 \text{ kJ/mol})] \\ &= -103.8 \text{ kJ} - (1 \text{ mol})(\Delta H_f^\circ \text{C}_3\text{H}_6(\text{g})) \\ (1 \text{ mol})(\Delta H_f^\circ \text{C}_3\text{H}_6(\text{g})) &= -103.8 \text{ kJ} + 124.2 \text{ kJ} \\ \Delta H_f^\circ \text{C}_3\text{H}_6(\text{g}) &= +20.4 \text{ kJ/mol} \end{aligned}$ <p>The enthalpy of formation of propene is 20.4 kJ/mol.</p>

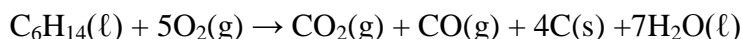
### Check Your Solution

Check to see that substitution of the data has been carried out correctly. The answer shows the same precision as the given data, one digit to the right of the decimal.

59. The complete combustion of hexane,  $\text{C}_6\text{H}_{14}(\ell)$ , can be represented by the following equation:



One possible reaction for the incomplete combustion of  $\text{C}_6\text{H}_{14}(\ell)$  is as follows:



Determine the amount of energy lost per mole of  $\text{C}_6\text{H}_{14}(\ell)$  when combustion is incomplete as written. The standard molar enthalpy of formation for  $\text{C}_6\text{H}_{14}(\ell)$  is  $-198.7 \text{ kJ/mol}$ .

### What Is Required?

You need to determine the difference in energy given off when combustion of hexane is incomplete compared to when it is complete.

### What Is Given?

You know the balanced thermochemical equation for the complete combustion of hexane. From this, you know the heat of combustion:  $-4163.2 \text{ kJ/mol}$

You have an equation for the incomplete combustion of hexane.

From Appendix B (student textbook page 743), you can obtain the molar enthalpies of formation:

$$\Delta H_f^\circ \text{ of } \text{C}_6\text{H}_{14}(\ell) = -198.7 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } \text{H}_2\text{O}(\ell) = -285.8 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ of } \text{CO}(\text{g}) = -110.5 \text{ kJ/mol}$$

### Plan Your Strategy

Use the available  $\Delta H_f^\circ$  data to substitute into the equation  $\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ , multiplying by the molar coefficients.

### Act on Your Strategy

$$\begin{aligned} \Delta H_r^\circ &= \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants}) \\ &= [(1 \text{ mol})(\Delta H_f^\circ \text{CO}_2(\text{g})) + (1 \text{ mol})(\Delta H_f^\circ \text{CO}(\text{g})) + (4 \text{ mol})(\Delta H_f^\circ \text{C}(\text{s})) \\ &\quad + (7 \text{ mol})(\Delta H_f^\circ \text{H}_2\text{O}(\ell))] - [(1 \text{ mol})(\Delta H_f^\circ \text{C}_6\text{H}_{14}(\ell)) + (5 \text{ mol})(\Delta H_f^\circ \text{O}_2(\text{g}))] \\ &= [(1 \text{ mol})(-393.5 \text{ kJ/mol}) + (1 \text{ mol})(-110.5 \text{ kJ/mol}) + (4 \text{ mol})(0 \text{ kJ/mol}) \\ &\quad + (7 \text{ mol})(-285.8 \text{ kJ/mol})] - [(1 \text{ mol})(-198.7 \text{ kJ/mol}) + (5 \text{ mol})(0 \text{ kJ/mol})] \\ &= -393.5 \text{ kJ} - 110.5 \text{ kJ} - 2000.6 + 198.7 \text{ kJ} \\ &= -2309.5 \text{ kJ} \end{aligned}$$

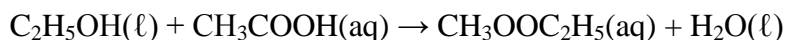


Plan Your Strategy	Act on Your Strategy
Determine the amount of energy lost per mole when combustion is incomplete by subtracting the enthalpy of reaction, $\Delta H_r^\circ$ , from the standard enthalpy of combustion, $\Delta H_{\text{comb}}^\circ$ .	$\begin{aligned}\text{Energy lost} &= \Delta H_{\text{comb}}^\circ - \Delta H_r^\circ \\ &= -4163.2 \text{ kJ/mol} - (-2305.9 \text{ kJ/mol}) \\ &= -1857.3 \text{ kJ/mol}\end{aligned}$

### Check Your Solution

The energy difference is very significant as expected for incomplete combustion. Check to see that substitution of the data has been carried out correctly. The answer shows the same precision as the given data, one digit to the right of the decimal.

60. Ethanol,  $\text{C}_2\text{H}_5\text{OH}(\ell)$ , reacts with a 1.00 mol/L solution of acetic acid to produce ethyl acetate, a solvent used to make protective coatings such as polyurethane.



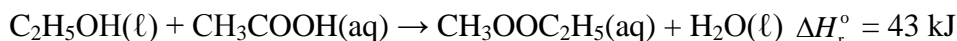
Calorimetry experiments determine that the enthalpy of reaction is 43 kJ/mol. Use this information and the enthalpy of formation data to calculate the enthalpy of formation of ethyl acetate. The enthalpy of formation for 1.00 mol/L  $\text{CH}_3\text{COOH}(\text{aq})$  is  $-486.0$  kJ/mol.

### What Is Required?

You have to find the enthalpy of formation for ethyl acetate.

### What Is Given?

You know the balanced chemical equation for the reaction and the enthalpy of reaction:



You know the molar enthalpy of formation of acetic acid:

$$\Delta H_f^\circ \text{ of } \text{CH}_3\text{COOH}(\text{aq}) = -486.0 \text{ kJ/mol}$$

From Appendix B (student textbook page 743), you can obtain the molar enthalpy of formation of ethanol:

$$\Delta H_f^\circ \text{ of } \text{C}_2\text{H}_5\text{OH}(\ell) = -277.6 \text{ kJ/mol}$$

Plan Your Strategy
Use the available $\Delta H_f^\circ$ data and $\Delta H_r^\circ$ data to substitute into the equation $\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ , multiplying by the molar coefficients. Then solve for $\Delta H_f^\circ$ of $\text{CH}_3\text{OOC}_2\text{H}_5(\text{aq})$ .
Act on Your Strategy
$\Delta H_r^\circ = \sum(n\Delta H_f^\circ \text{ products}) - \sum(n\Delta H_f^\circ \text{ reactants})$ $43 \text{ kJ} = [(1 \text{ mol})(\Delta H_f^\circ \text{CH}_3\text{OOC}_2\text{H}_5(\text{aq})) + (1 \text{ mol})(\Delta H_f^\circ \text{H}_2\text{O}(\ell))] - [(1 \text{ mol})(\Delta H_f^\circ \text{C}_2\text{H}_5\text{OH}(\ell)) + (1 \text{ mol})(\Delta H_f^\circ \text{CH}_3\text{COOH}(\text{aq}))]$ $= [(1 \text{ mol})(\Delta H_f^\circ \text{CH}_3\text{OOC}_2\text{H}_5(\text{aq})) + (1 \cancel{\text{mol}})(-241.8 \text{ kJ}/\cancel{\text{mol}})] - [(1 \text{ mol})(-277.6 \text{ kJ}/\cancel{\text{mol}}) + (1 \cancel{\text{mol}})(-486.0 \text{ kJ}/\cancel{\text{mol}})]$ $= (1 \text{ mol})(\Delta H_f^\circ \text{CH}_3\text{OOC}_2\text{H}_5(\text{aq})) + 521.8 \text{ kJ}$ $43 \text{ kJ} - 521.8 \text{ kJ} = (1 \text{ mol})(\Delta H_f^\circ \text{CH}_3\text{OOC}_2\text{H}_5(\text{aq}))$ $\Delta H_f^\circ \text{CH}_3\text{OOC}_2\text{H}_5(\text{aq}) = -478.8 \text{ kJ/mol}$

### Check Your Solution

Check to see that substitution of the data has been carried out correctly. The answer shows the same precision as the given data, one digit to the right of the decimal.

## Determining the Efficiency of a Propane Barbecue

**(Student textbook page 327)**

61. Using the data for the molar enthalpy of combustion of butane from **Table 5.4** (student textbook page 295), determine the efficiency of a lighter as a heating device if 0.70 g of butane is required to increase the temperature of a 250.0 g stainless steel spoon ( $c = 0.503 \text{ J/g}\cdot^\circ\text{C}$ ) by  $45.0^\circ\text{C}$ .

**What Is Required?**

You need to determine the efficiency of a butane lighter.

**What Is Given?**

You know the mass,  $m$ , of butane: 0.70 g

You know the mass,  $m$ , of stainless steel: 250.0 g

You know the specific heat capacity,  $c$ , of stainless steel:  $0.503 \text{ J/g}\cdot^\circ\text{C}$

You know the change in temperature change,  $\Delta T$ :  $45.0^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Determine the amount in moles, $n$ , of butane, using the mass and molar mass of butane, and the formula $n = \frac{m}{M}$ .	$n = \frac{m}{M}$ $n_{\text{C}_4\text{H}_{10}} = \frac{0.70 \cancel{\text{g}}}{58.14 \cancel{\text{g}}/\text{mol}}$ $= 0.012\,039 \text{ mol}$
Use the formula $\Delta H = n\Delta H_{\text{comb}}$ to determine the enthalpy change, $\Delta H$ , of the butane that is burned.	$\Delta H = n\Delta H_{\text{comb}}$ $= (0.012039 \cancel{\text{mol}})(-2877.6\text{kJ}/\cancel{\text{mol}})$ $= -34.643\,4 \text{ kJ}$
Since the energy input is from the burning of butane, change the sign of $\Delta H$ to determine the energy input.	$\text{energy input} = -\Delta H$ $= -(-34.643\,4 \text{ kJ})$ $= 34.643\,4 \text{ kJ}$
Use the formula $Q = mc\Delta T$ to calculate the energy output, $Q$ , for the stainless steel.	$Q = mc\Delta T$ $= (250 \cancel{\text{g}})(0.503\text{J}/\cancel{\text{g}}\cdot^\circ\text{C})(45.0\cancel{^\circ\text{C}})$ $= 5658.75 \text{ J}$ $= 5.658\,8 \text{ kJ}$

Calculate the efficiency of the lighter.	$\text{Efficiency} = \frac{\text{energy output}}{\text{energy input}} \times 100\%$ $= \frac{5.658 \text{ 8 } \cancel{\text{kJ}}}{34.643 \text{ 4 } \cancel{\text{kJ}}} \times 100\%$ $= 16.334\%$ <p>The efficiency is 16%.</p>
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### Check Your Solution

The efficiency is expected to be very low since there would be a great deal of heat loss when a lighter is use to heat a spoon. The answer seems reasonable and has the correct number of significant digits.

62. A solid camping fuel has a stated energy content of 50.0 kJ/g. Determine its efficiency if a 2.50 g piece of the fuel was required to increase the temperature of 500.0 g of soup ( $c = 3.77 \text{ J/g}\cdot^\circ\text{C}$ ) in a 50.0 g aluminum pot by  $45.0^\circ\text{C}$ .

### What Is Required?

You need to determine the efficiency of a camping fuel to heat a sample of soup in an aluminum pot.

### What Is Given?

You know the mass of fuel:  $m_{\text{fuel}} = 2.50 \text{ g}$

You know the mass of soup:  $m_{\text{soup}} = 500.0 \text{ g}$

You know the specific heat capacity of the soup:  $c_{\text{soup}} = 3.77 \text{ J/g}\cdot^\circ\text{C}$

You know the mass,  $m$ , of the aluminum pot:  $m_{\text{Al}} = 50.0 \text{ g}$

You know the specific heat capacity of Al:  $c_{\text{Al}} = 0.897 \text{ J/g}\cdot^\circ\text{C}$

You know the change in temperature:  $\Delta T = 45.0^\circ\text{C}$

Plan Your Strategy	Act on Your Strategy
Determine the energy input from the burning fuel using: energy input = $m \times \text{energy content}$ (per gram).	$\begin{aligned}\text{energy input} &= m \times \text{energy content} \\ &= 2.50 \text{ g} \times 50.0 \text{ kJ/g} \\ &= 125 \text{ kJ}\end{aligned}$
Determine the energy output, $Q_{\text{total}}$ , absorbed by the soup and the aluminum pot using the formula $Q = mc\Delta T$ .	$\begin{aligned}Q_{\text{total}} &= Q_{\text{soup}} + Q_{\text{Al}} \\ &= (mc\Delta T)_{\text{soup}} + (mc\Delta T)_{\text{Al}} \\ &= (500.0 \text{ g} \times 3.77 \text{ J/g}\cdot^\circ\text{C} \times 45.0^\circ\text{C}) \\ &\quad + (50.0 \text{ g} \times 0.897 \text{ J/g}\cdot^\circ\text{C} \times 45.0^\circ\text{C}) \\ &= 84\,825 \text{ J} + 2018.25 \text{ J} \\ &= 86\,843.25 \text{ J} \\ &= 86.843 \text{ kJ}\end{aligned}$
Calculate the efficiency of the burning camping fuel.	$\begin{aligned}\text{Efficiency} &= \frac{\text{energy output}}{\text{energy input}} \times 100\% \\ &= \frac{86.843 \text{ kJ}}{125 \text{ kJ}} \times 100\% \\ &= 69.4747\% \\ &= 69.5\%\end{aligned}$ <p>The efficiency is 69.5%.</p>

### Check Your Solution

The efficiency is less than 100%, which is expected since there will be heat loss to the environment. The answer seems reasonable and shows the correct number of significant digits.

63. Determine the efficiency of a heating device that burns methanol,  $\text{CH}_3\text{OH}(\ell)$ , given the following information:

**Data for Determining the Efficiency of a Methanol-burning Heater**

Quantity Being Measured	Data
Initial mass of burner	38.37 g
Final mass of burner	36.92 g
Mass of aluminum can	257.36 g
Mass of aluminum can and water	437.26 g
Initial temperature of water	10.45°C
Final temperature of water	23.36°C
$\Delta H^\circ_{\text{comb}}(\text{CH}_3\text{OH})$	-726.1 kJ/mol

**What Is Required?**

You have to calculate the efficiency of a methanol burner.

**What Is Given?**

The information is given in the data table in the problem.

Plan Your Strategy	Act on Your Strategy
Determine the mass of methanol burned.	$m_{\text{CH}_3\text{OH}} = m_{\text{initial}} - m_{\text{final}}$ $= 38.37 \text{ g} - 36.92 \text{ g}$ $= 1.45 \text{ g}$
Use the mass and molar mass of methanol and the formula $n = \frac{m}{M}$ to calculate the amount in moles, $n$ , of methanol burned.	$n = \frac{m}{M}$ $n_{\text{CH}_3\text{OH}} = \frac{1.45 \text{ g}}{32.05 \text{ g/mol}}$ $= 0.04524 \text{ mol}$
Use the formula $\Delta H = n\Delta H^\circ_{\text{comb}}$ to calculate the energy input, $\Delta H$ , from the burning of methanol.	$\Delta H = n\Delta H^\circ_{\text{comb}}$ $= (0.04524 \text{ mol})(-726.1 \text{ kJ/mol})$ $= -32.840 \text{ kJ}$ <p>The energy input is 32.850 kJ.</p>
Calculate the mass, $m$ , of water in the can.	$m = m_{\text{can + water}} - m_{\text{can}}$ $= 437.26 \text{ g} - 257.36 \text{ g}$ $= 179.9 \text{ g}$
Determine the change in temperature using the formula $\Delta T = T_{\text{final}} - T_{\text{initial}}$ .	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 23.36^\circ\text{C} - 10.4^\circ\text{C}$ $= 12.91^\circ\text{C}$

Determine the energy output, $Q_{\text{total}}$ , absorbed by the aluminum can and water using the formula $Q = mc\Delta T$ .	$  \begin{aligned}  Q_{\text{total}} &= Q_{\text{Al}} + Q_{\text{H}_2\text{O}} \\  &= (mc\Delta T)_{\text{Al}} + (mc\Delta T)_{\text{H}_2\text{O}} \\  &= (257.36 \cancel{\text{ g}} \times 0.897 \text{ J} / \cancel{\text{ g} \cdot ^\circ\text{C}} \times 12.91 \cancel{^\circ\text{C}}) \\  &\quad + (179.9 \cancel{\text{ g}} \times 4.19 \text{ J} / \cancel{\text{ g} \cdot ^\circ\text{C}} \times 12.91 \cancel{^\circ\text{C}}) \\  &= 2980.3 \text{ J} + 9731.3 \text{ J} \\  &= 12\,711.61 \text{ J} \\  &= 12.7116 \text{ kJ}  \end{aligned}  $
Calculate the efficiency of the heating device.	$  \begin{aligned}  \text{Efficiency} &= \frac{\text{energy output}}{\text{energy input}} \times 100\% \\  &= \frac{12.7116 \cancel{\text{ kJ}}}{32.84 \cancel{\text{ kJ}}} \times 100\% \\  &= 38.695\% \\  &= 38.7\%  \end{aligned}  $

### Check Your Solution

The efficiency is much less than 100%, which is expected since there are heat losses to the surrounding air with this calorimeter set-up. The answer has the correct number of significant digits.

64. What mass of pentane,  $\text{C}_5\text{H}_{12}(\text{g})$ , would have to be burned in an open system to heat  $2.50 \times 10^2 \text{ g}$  of hot chocolate ( $c = 3.59 \text{ J/g}\cdot^\circ\text{C}$ ) from  $20.0^\circ\text{C}$  to  $39.8^\circ\text{C}$  if the energy conversion is 45.0% efficient?

### What Is Required?

You need to calculate the mass of pentane that is necessary to heat hot chocolate in an open system if the process is 45.0% efficient.

### What Is Given?

You know the mass,  $m$ , of the hot chocolate:  $2.50 \times 10^2 \text{ g}$

You know the specific heat capacity,  $c$ , of the hot chocolate:  $3.59 \text{ J/g}\cdot^\circ\text{C}$

You know the initial temperature,  $T_{\text{initial}}$ :  $20.0^\circ\text{C}$

You know the final temperature,  $T_{\text{final}}$ :  $39.8^\circ\text{C}$

You know the efficiency of the process: 45.0%

Plan Your Strategy	Act on Your Strategy
Determine the change in temperature using the formula $\Delta T = T_{\text{final}} - T_{\text{initial}}$ .	$\begin{aligned}\Delta T &= T_{\text{final}} - T_{\text{initial}} \\ &= 39.8^\circ\text{C} - 20.0^\circ\text{C} \\ &= 19.8^\circ\text{C}\end{aligned}$
Use the formula $Q = mc\Delta T$ to calculate the energy output, $Q$ , that is absorbed by the hot chocolate.	$\begin{aligned}Q &= mc\Delta T \\ &= (2.50 \times 10^2 \text{ g})(3.59 \text{ J/g}\cdot^\circ\text{C})(19.8^\circ\text{C}) \\ &= 17\,770.5 \text{ J} \\ &= 17.770\,5 \text{ kJ}\end{aligned}$
Use the formula for efficiency to calculate the energy input from the burning of pentane.	$\begin{aligned}\text{Efficiency} &= \frac{\text{energy output}}{\text{energy input}} \times 100\% \\ 45\% &= \frac{17.770\,5 \text{ kJ}}{\text{energy input}} \times 100\% \\ \text{energy input} &= 17.770\,5 \text{ kJ} \times \frac{100\%}{45\%} \\ &= 39.4888 \text{ kJ}\end{aligned}$
Since the energy input is from the burning of pentane, change the sign of the energy input to determine $\Delta H$ of pentane.	$\begin{aligned}\Delta H &= -(\text{energy input}) \\ &= -39.4888 \text{ kJ}\end{aligned}$
Use the formula $\Delta H_{\text{comb}}^\circ = \frac{\Delta H}{n}$ to calculate the amount in moles, $n$ , of pentane burned.	$\begin{aligned}\Delta H_{\text{comb}}^\circ &= \frac{\Delta H}{n} \\ -3509.0 \text{ kJ/mol} &= \frac{-39.4888 \text{ kJ}}{n} \\ n &= \frac{-39.4888 \text{ kJ}}{-3509.0 \text{ kJ/mol}} \\ &= 0.011\,253\,6 \text{ mol}\end{aligned}$



Using the amount in moles and molar mass of pentane, $\text{C}_5\text{H}_{12}(\text{g})$ , use the formula $m = nM$ to determine the mass, $m$ , of pentane.	$m = nM$ $m_{\text{C}_5\text{H}_{12}} = (0.011\,253\,6\,\cancel{\text{mol}})(72.17\,\text{g}/\cancel{\text{mol}})$ $= 0.812172\,\text{g}$ <p>The mass of pentane is 0.812 g.</p>
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### Check Your Solution

The units and the number of significant digits are correct. For heating that is less than 50% efficient, the mass of pentane required seems reasonable since pentane has a fairly high heat of combustion.

65. Heat from burning a fuel is absorbed by a copper calorimeter. The calorimeter has a mass of 81.34 g and contains 200.00 g of water initially at 21.00°C.
- How much heat was absorbed by the calorimeter and the water if the final temperature of the water is 36.40°C? Refer to **Table 5.1** for the required specific heat capacities.
  - The absorption of heat energy by the system is 45.00% efficient. How much heat was lost to the surroundings?

### What Is Required?

You need to calculate the heat absorbed by a calorimeter and water as a fuel burns.

### What Is Given?

You know the mass of the calorimeter:  $m_{\text{calorimeter}} = 81.34 \text{ g}$

You know the specific heat capacity of copper:  $c_{\text{Cu}} = 0.385 \text{ J/g}\cdot^{\circ}\text{C}$

You know the mass of water:  $m_{\text{H}_2\text{O}} = 200.00 \text{ g}$

You know the specific heat capacity of water:  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^{\circ}\text{C}$

You know the initial temperature:  $T_{\text{initial}} = 21.0^{\circ}\text{C}$

You know the final temperature:  $T_{\text{final}} = 36.4^{\circ}\text{C}$

You know that  $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ .

Plan Your Strategy	Act on Your Strategy
Determine the change in temperature using the formula $\Delta T = T_{\text{final}} - T_{\text{initial}}$ .	$\begin{aligned}\Delta T &= T_{\text{final}} - T_{\text{initial}} \\ &= 36.40^{\circ}\text{C} - 21.00^{\circ}\text{C} \\ &= 15.40^{\circ}\text{C}\end{aligned}$
Use the formula $Q = mc\Delta T$ to calculate the energy output, $Q_{\text{total}}$ , that is absorbed by the copper calorimeter and water.	$\begin{aligned}Q_{\text{total}} &= Q_{\text{Cu}} + Q_{\text{H}_2\text{O}} \\ &= (mc\Delta T)_{\text{Cu}} + (mc\Delta T)_{\text{H}_2\text{O}} \\ &= (81.34 \cancel{\text{g}} \times 0.385 \text{ J} / \cancel{\text{g}} \cdot \cancel{^{\circ}\text{C}} \times 15.40 \cancel{^{\circ}\text{C}}) \\ &\quad + (200.00 \cancel{\text{g}} \times 4.19 \text{ J} / \cancel{\text{g}} \cdot \cancel{^{\circ}\text{C}} \times 15.40 \cancel{^{\circ}\text{C}}) \\ &= 482.265 \text{ J} + 12\,905.2 \text{ J} \\ &= 13\,387.465 \text{ J} \\ &= 13.387\,781 \text{ kJ} \\ &= 13.4 \text{ kJ}\end{aligned}$

### Check Your Solution

The temperature of the calorimeter and water increases so heat was gained, and the sign of the heat gain is positive. The answer shows the correct number of significant digits and seems reasonable.

**b. What Is Required?**

You need to determine the heat lost to the surroundings.

**What Is Given?**

You know the efficiency of the system is 45.00%.

<b>Plan Your Strategy</b>	<b>Act on Your Strategy</b>
Determine the percent of heat lost to the surroundings.	If the system is 45.00% efficient, 55.00% of the heat is lost to the surroundings.
Calculate the amount of energy that this percent represents of the energy given off from the burning fuel.	heat lost = (0.5500)(13.387 781 kJ) = 7363.2795 kJ  The loss is 7363 kJ = $7.36 \times 10^3$ kJ.

**Check Your Solution**

Check that the data was multiplied correctly. The answer has the correct number of significant digits.

66. A furnace outputs  $5.000 \times 10^4$  kJ of heat energy to an empty room. The following information is given:

dimensions of room:  $10.0 \text{ m} \times 10.0 \text{ m} \times 10.0 \text{ m}$

density of air:  $1.290 \text{ kg/m}^3$

specific heat capacity of air:  $1.01 \text{ J/g}\cdot^\circ\text{C}$

initial air temperature:  $12.0^\circ\text{C}$

final air temperature:  $26.0^\circ\text{C}$

- Calculate the quantity of heat gained by the air in this room.
- Determine how much heat was lost from the room, and calculate an efficiency rating for the retention of heat for this room.

### What Is Required?

- You need to calculate the heat that is gained by the air in an empty room.
- You need to determine the quantity of heat lost and efficiency rating for the retention of heat.

### What Is Given?

You know the dimensions of the room:  $l = w = h = 10.0 \text{ m}$

You know the density of air:  $d = 1.290 \text{ kg/m}^3$

You know the specific heat capacity of air from **Table 5.1**:  $c_{\text{air}} = 1.01 \text{ J/g}\cdot^\circ\text{C}$

You know the initial temperature,:  $T_{\text{initial}} = 12.0^\circ\text{C}$

You know the final temperature,:  $T_{\text{final}} = 26.0^\circ\text{C}$

You know the energy output from the furnace:  $5.000 \times 10^4 \text{ kJ}$

Plan Your Strategy	Act on Your Strategy
Use the dimensions of the room to determine the volume, $V$ , of air.	$V = lwh$ $= (10.0 \text{ m})(10.0 \text{ m})(10.0 \text{ m})$ $= 1 \times 10^3 \text{ m}^3$
Use the formula $m = d \times V$ to calculate the mass of air.	$m = dV$ $= (1.290 \text{ kg/m}^3)(1 \times 10^3 \text{ m}^3)$ $= 1.290 \times 10^3 \text{ kg}$ $= 1.290 \times 10^6 \text{ g}$
Determine the change in temperature using the formula $\Delta T = T_{\text{final}} - T_{\text{initial}}$ .	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 26.0^\circ\text{C} - 12.0^\circ\text{C}$ $= 14.0^\circ\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the energy output, $Q$ , that is absorbed by the air.	$Q = mc\Delta T$ $= (1.290 \times 10^6 \text{ g})(1.01 \text{ J/g}\cdot^\circ\text{C})(14.0^\circ\text{C})$ $= 1.824\,060\,0 \times 10^7 \text{ J}$ $= 1.824\,060\,0 \times 10^4 \text{ kJ}$  The quantity of heat retained in the room is $1.82 \times 10^4 \text{ kJ}$ .

Use the output of energy from the furnace and the heat gained by the air in the room to determine the heat loss.	$\begin{aligned}\text{heat loss} &= \text{energy input} - \text{output retained} \\ &= (5.000 \times 10^4 \text{ kJ}) - (1.82 \times 10^4 \text{ kJ}) \\ &= 3.18 \times 10^4 \text{ kJ}\end{aligned}$
Use the formula for efficiency to determine the efficiency rating.	$\begin{aligned}\text{Efficiency} &= \frac{\text{energy output}}{\text{energy input}} \times 100\% \\ &= \frac{1.82 \times 10^4 \cancel{\text{ kJ}}}{5.000 \times 10^4 \cancel{\text{ kJ}}} \times 100\% \\ &= 36.4\%\end{aligned}$ <p>The room has an efficiency rating of 36.4% to retain heat.</p>

### Check Your Solution

- a. The answer is large, but you must consider that this temperature change would occur over a period of time. The answer has the correct number of significant digits.
- b. The efficiency is low, indicating a large heat loss in the room. This can be attributed to insulation factors. The answer shows the correct number of significant digits.

67. A steel steam boiler has a mass of  $1.50 \times 10^3$  kg and contains 255.00 kg of water at  $35.00^\circ\text{C}$ . Heat absorption by the boiler is 75.45% efficient. How much heat is required to increase the temperature of the system to the boiling point of water? The specific heat capacity of steel is  $0.488 \text{ J/g}\cdot^\circ\text{C}$ . The answer may be expressed in megajoules ( $1 \text{ MJ} = 1 \times 10^6 \text{ J}$ ).

### What Is Required?

You need to determine the amount of heat required to raise the temperature of water to its boiling point.

### What Is Given?

You know the mass,  $m$ , of the steel boiler:  $1.50 \times 10^3$  kg

You know the specific heat capacity,  $c$ , of steel:  $0.488 \text{ J/g}\cdot^\circ\text{C}$

You know the mass,  $m$ , of water: 255.00 kg

You know the specific heat capacity,  $c$ , of water:  $4.19 \text{ J/g}\cdot^\circ\text{C}$

You know the initial temperature,  $T_{\text{initial}}$ :  $35.00^\circ\text{C}$

You know the final temperature,  $T_{\text{final}}$ :  $100.00^\circ\text{C}$

You know the efficiency of the system: 75.45%

Plan Your Strategy	Act on Your Strategy
Determine the change in temperature using the formula $\Delta T = T_{\text{final}} - T_{\text{initial}}$ .	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 100.00^\circ\text{C} - 35.00^\circ\text{C}$ $= 65.00^\circ\text{C}$
Use the formula $Q = mc\Delta T$ to calculate the energy output, $Q_{\text{total}}$ , that is absorbed by the boiler and water.	$Q_{\text{total}} = Q_{\text{steel}} + Q_{\text{H}_2\text{O}}$ $= (mc\Delta T)_{\text{steel}} + (mc\Delta T)_{\text{H}_2\text{O}}$ $= (1\,500\,000 \text{ g} \times 0.488 \text{ J/g}\cdot^\circ\text{C} \times 65.00^\circ\text{C})$ $+ (255\,000.00 \text{ g} \times 4.19 \text{ J/g}\cdot^\circ\text{C} \times 65.00^\circ\text{C})$ $= 47\,580\,000 \text{ J} + 65\,449\,250 \text{ J}$ $= 117\,029\,250 \text{ J}$
Use the formula for efficiency to calculate the energy input from the heating system.	$\text{Efficiency} = \frac{\text{energy output}}{\text{energy input}} \times 100\%$ $75.45\% = \frac{117\,029\,250 \text{ J}}{\text{energy input}} \times 100\%$ $\text{energy input} = 117\,029\,250 \text{ J} \times \frac{100\%}{75.45\%}$ $= 155\,108\,349.9 \text{ J}$ $= 155.1 \text{ MJ}$ <p>The energy that must be added to the boiler system is 155 MJ.</p>

### Check Your Solution

The amount of energy is very large, which is indicative of the energy demand of a water boiler. The answer has the correct number of significant digits.

68. A flame calorimeter is used to determine the heat content of wood. The aluminum can in the apparatus has a mass of 180.51 g and holds 1.000 kg of water initially at 15.6°C. After 5.50 g of wood is burned, the water temperature increases to 20.7°C.
- Calculate the energy output that is absorbed by the calorimeter and the water.
  - An estimate of the recoverable heat content for the wood is 6.5 kJ/g. How much heat energy is estimated to be available in the sample of wood?
  - Determine the efficiency of the system in measuring the heat content of the wood.

**a. What Is Required?**

You need to determine the energy absorbed by a calorimeter and water.

**What Is Given?**

You know the mass of the aluminum calorimeter:  $m_{\text{Al}} = 180.51 \text{ g}$

You know the specific heat capacity of aluminum:  $c_{\text{Al}} = 0.897 \text{ J/g}\cdot^{\circ}\text{C}$

You know the mass of water:  $m_{\text{H}_2\text{O}} = 1.000 \text{ kg}$

You know the specific heat capacity of water:  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^{\circ}\text{C}$

You know the initial temperature:  $T_{\text{initial}} = 15.6^{\circ}\text{C}$

You know the final temperature:  $T_{\text{final}} = 20.7^{\circ}\text{C}$

Plan Your Strategy	Act on Your Strategy
Determine the change in temperature using the formula $\Delta T = T_{\text{final}} - T_{\text{initial}}$ .	$\begin{aligned}\Delta T &= T_{\text{final}} - T_{\text{initial}} \\ &= 20.7^{\circ}\text{C} - 15.6^{\circ}\text{C} \\ &= 5.1^{\circ}\text{C}\end{aligned}$
Use the formula $Q = mc\Delta T$ to calculate the energy output, $Q_{\text{total}}$ , that is absorbed by the aluminum calorimeter and water.	$\begin{aligned}Q_{\text{total}} &= Q_{\text{Al}} + Q_{\text{H}_2\text{O}} \\ &= (mc\Delta T)_{\text{Al}} + (mc\Delta T)_{\text{H}_2\text{O}} \\ &= (180.51 \cancel{\text{g}} \times 0.897 \text{ J} / \cancel{\text{g}}\cdot^{\circ}\text{C} \times 5.1\cancel{^{\circ}\text{C}}) \\ &\quad + (1000 \cancel{\text{g}} \times 4.19 \text{ J} / \cancel{\text{g}}\cdot^{\circ}\text{C} \times 5.1\cancel{^{\circ}\text{C}}) \\ &= 825.779 \text{ J} + 21\,369.00 \text{ J} \\ &= 22\,194.7791 \text{ J} \\ &= 2.2 \times 10^4 \text{ J}\end{aligned}$

**b. What is Required?**

You need to calculate the energy content of wood.

**What Is Given?**

You know the heat content of wood: 6.5 kJ/g

You know the mass of wood burned: 5.50 g

Plan Your Strategy	Act on Your Strategy
Use the heat content and mass of wood to determine the heat input from the burning wood.	$\begin{aligned}\text{heat input} &= m \times \text{heat content} \\ &= (5.5 \cancel{\text{ g}})(6.5 \text{ kJ}/\cancel{\text{ g}}) \\ &= 35.75 \text{ kJ} \\ &= 36 \text{ kJ}\end{aligned}$

**c. What is Required?**

You need to determine the efficiency of the system using this wood.

**What is Given?**

You know the energy output is 22194.7791 J

You know the energy input is 35.75 kJ

Plan Your Strategy	Act on Your Strategy
Convert the energy output to kJ.	$22194.7791 \text{ J} = 22.194 \text{ 78 kJ}$
Use the formula for efficiency to calculate the energy input from the heating system.	$\begin{aligned}\text{Efficiency} &= \frac{\text{energy absorbed}}{\text{energy available}} \times 100\% \\ &= \frac{22.194 \text{ 78} \cancel{\text{ kJ}}}{35.75 \cancel{\text{ kJ}}} \times 100\% \\ &= 62.08\%\end{aligned}$ <p>The efficiency of the calorimeter in determining the heat content of wood is 62%.</p>

**Check Your Solution**

Check that data has been correctly substituted into the formulas. The answers have the correct number of significant digits and seem reasonable.



69. Refer to **Tables 5.3** (student textbook page 290) and **5.4** (student textbook page 295) to obtain data necessary for these questions.
- What mass of methane must burn to heat 60.00 g of water from 25.0°C to its boiling point and then turn it into steam at 100.0°C? (**Hint:** Two steps are involved: heating of the water to 100.0°C, and then change of the liquid water to a vapour at 100.0°C).
  - If the transfer of heat is 56.5% efficient, what quantity of heat must come from the burning of methane?
  - What mass of methane must be burned at this efficiency rating?

### What Is Required?

- You need to calculate the mass of methane required to warm water to steam.
- You need to determine the energy input from the burning of methane at a given efficiency.
- You need to find the mass of methane required at the given efficiency.

### What Is Given?

You know the mass,  $m$ , of water: 60.00 g

You know the initial temperature:  $T_{\text{initial}} = 25.0^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 100.0^\circ\text{C}$

You know the heat of vapourization of water: 40.7 kJ/mol

You know the efficiency of system: 56.5%

For 100% efficiency you know that  $\Delta E_{\text{system}} = -\Delta E_{\text{surroundings}}$ . Therefore, the heat given off by burning methane is equal to the heat gained by water.

### a. Mass of methane required to warm water to steam

Plan Your Strategy	Act on Your Strategy
Determine the change in temperature using the formula $\Delta T = T_{\text{final}} - T_{\text{initial}}$ .	$\Delta T = T_{\text{final}} - T_{\text{initial}}$ $= 100.0^\circ\text{C} - 25.0^\circ\text{C}$ $= 75.0^\circ\text{C}$
Use mass and molar mass of water and the formula $n = \frac{m}{M}$ to calculate the amount in moles, $n$ , of water.	$n = \frac{m}{M}$ $n_{\text{H}_2\text{O}} = \frac{60.0 \text{ g}}{18.02 \text{ g/mol}}$ $= 3.329 \text{ mol}$
Use the formula $Q = mc\Delta T + n\Delta H_{\text{vap}}$ to calculate the heat, $Q$ , absorbed by the water.	$Q = mc\Delta T + n\Delta H_{\text{vap}}$ $= (60.0 \text{ g}) \left( 4.19 \frac{\text{J}}{\text{g}^\circ\text{C}} \right) (75.0^\circ\text{C})$ $+ (3.329 \text{ mol}) \left( 40.7 \frac{\text{kJ}}{\text{mol}} \right)$ $= 18\,855 \text{ J} + 135.49 \text{ kJ}$ $= 18.855 \text{ kJ} + 135.49 \text{ kJ}$ $= 154.3453 \text{ kJ}$

Change the sign of $Q$ to find the amount of thermal energy, $\Delta H$ , given off during the burning of methane.	$\Delta H = -Q$ $= -154.3453 \text{ kJ}$
Use the standard enthalpy of combustion of methane and the formula $\Delta H_{\text{comb}}^{\circ} = \frac{\Delta H}{n}$ to calculate the amount in moles, $n$ , of methane.	$\Delta H_{\text{comb}}^{\circ} = \frac{\Delta H}{n}$ $-890.8 \text{ kJ/mol} = \frac{-154.3453 \text{ kJ}}{n_{\text{CH}_4}}$ $n_{\text{CH}_4} = \frac{-154.3453 \cancel{\text{ kJ}}}{-890.8 \cancel{\text{ kJ}}/\text{mol}}$ $= 0.173\,265\,9 \text{ mol}$
Use the amount in moles and molar mass of methane, and the formula $m = nM$ to calculate the mass, $m$ , of methane.	$m = nM$ $m_{\text{CH}_4} = nM$ $= (0.173\,265\,9 \cancel{\text{ mol}})(16.05 \text{ g}/\cancel{\text{ mol}})$ $= 2.780\,92 \text{ g}$ $= 2.78 \text{ g}$ <p>The mass of methane that must be burned is 2.78 g.</p>

**b. Quantity of heat from burning methane**

Plan Your Strategy	Act on Your Strategy
Use the formula for efficiency to determine the energy input required from the burning of methane.	$\text{Efficiency} = \frac{\text{energy output}}{\text{energy input}} \times 100\%$ $56.5\% = \frac{154.3453 \text{ J}}{\text{energy input}} \times 100\%$ $\text{energy input} = 154.3453 \text{ kJ} \times \frac{100\%}{56.5\%}$ $= 273.177 \text{ 52 kJ}$ <p>The energy that must come from the burning is 273 kJ.</p>

**c. Mass of methane required for given efficiency**

Plan Your Strategy	Act on Your Strategy
Calculate the mass, $m$ , of methane required at the given efficiency.	$m = 2.781 \text{ 37 g} \times \frac{100\%}{56.5\%}$ $= 4.92 \text{ g}$ <p>The mass of methane required is 4.92 g.</p>

**Check Your Solution**

A greater mass of methane must be burned since the system is less than 100% efficient. The answers have the correct number of significant digits and seem reasonable.

70. Ethanol,  $\text{C}_2\text{H}_5\text{OH}(\ell)$ , is used as a fuel in an alcohol burner. A glass beaker that has a mass of 386.00 g and a specific heat capacity of  $0.880 \text{ J/g}\cdot^\circ\text{C}$  is filled with 125.00 g of water initially at  $15.5^\circ\text{C}$ . The beaker of water is placed next to the burner. When 1.24 g of ethanol burns, the temperature of the water in the beaker increases to  $48.7^\circ\text{C}$ . What is the efficiency of the transfer of heat to the beaker and the water? The enthalpy of combustion of ethanol is  $1366.8 \text{ kJ/mol}$ .

### What Is Required?

You need to calculate the efficiency of the system when ethanol heats water.

### What Is Given?

You know the mass of the beaker:  $m_{\text{beaker}} = 386.00 \text{ g}$

You know the specific heat capacity of the beaker:  $c_{\text{beaker}} = 0.880 \text{ J/g}\cdot^\circ\text{C}$

You know the mass of water:  $m_{\text{H}_2\text{O}} = 125.00 \text{ g}$

You know the specific heat capacity of water:  $c_{\text{H}_2\text{O}} = 4.19 \text{ J/g}\cdot^\circ\text{C}$

You know the initial temperature:  $T_{\text{initial}} = 15.5^\circ\text{C}$

You know the final temperature:  $T_{\text{final}} = 48.7^\circ\text{C}$

You know the mass of ethanol burned:  $m_{\text{ethanol}} = 1.24 \text{ g}$

You know the heat of combustion for ethanol:  $\Delta H_{\text{comb}}^\circ = -1366.8 \text{ kJ/mol}$

Plan Your Strategy	Act on Your Strategy
Determine the change in temperature using the formula $\Delta T = T_{\text{final}} - T_{\text{initial}}$ .	$\begin{aligned}\Delta T &= T_{\text{final}} - T_{\text{initial}} \\ &= 48.7^\circ\text{C} - 15.5^\circ\text{C} \\ &= 33.2^\circ\text{C}\end{aligned}$
Use the formula $Q = mc\Delta T$ to determine the energy output, $Q$ , that is absorbed by the beaker and water.	$\begin{aligned}Q &= mc\Delta T \\ &= (mc\Delta T)_{\text{glass}} + (mc\Delta T)_{\text{water}} \\ &= (386.00 \cancel{\text{g}} \times 0.880 \text{ J} / \cancel{\text{g}}\cdot^\circ\text{C} \times 33.2\cancel{^\circ\text{C}}) \\ &\quad + (125.00 \cancel{\text{g}} \times 4.19 \text{ J} / \cancel{\text{g}}\cdot^\circ\text{C} \times 33.2\cancel{^\circ\text{C}}) \\ &= 11\,277.376 \text{ J} + 17\,388.500 \text{ J} \\ &= 28\,665.876 \text{ J} \\ &= 28.665\,87 \text{ kJ}\end{aligned}$
Use mass and molar mass of ethanol and the formula $n = \frac{m}{M}$ to calculate the amount in moles, $n$ , of ethanol.	$\begin{aligned}n &= \frac{m}{M} \\ n_{\text{C}_2\text{H}_5\text{OH}} &= \frac{1.24 \cancel{\text{g}}}{46.08 \cancel{\text{g}}/\text{mol}} \\ &= 0.026\,91 \text{ mol}\end{aligned}$

Use the formula $\Delta H = n\Delta H_{\text{comb}}^{\circ}$ to determine the energy input, $\Delta H$ , from the burning of ethanol.	$\Delta H = n\Delta H_{\text{comb}}^{\circ}$ $= (0.02691 \text{ mol})(-1366.8 \text{ kJ/mol})$ $= -36.7805 \text{ kJ}$
Use the formula for efficiency to calculate the efficiency of the system.	<p>Efficiency</p> $= \frac{\text{Energy output (heat absorbed by glass and water)}}{\text{Energy input (heat released during burning)}} \times 100\%$ $= \frac{28.66587 \text{ kJ}}{36.7805 \text{ kJ}} \times 100\%$ $= 77.93768\%$ $= 77.9\%$ <p>The transfer of heat has an efficiency of 77.9%.</p>

### Check Your Solution

The efficiency is less than 100% indicating that there is heat loss in this system. This is reasonable for this experimental set-up. The answer has the correct number of significant digits.

## Solutions to Practice Problems in Chapter 6 Rates of Reaction

### Calculating an Average Reaction Rate

#### Determining Reaction Rates in Terms of Products and Reactants

(Student textbook pages 360-1)

1. In the reaction  $A + 2B \rightarrow 3C + 4D$ , the initial concentration of A was 0.0415 mol/L, and after 14.7 min the concentration of A was 0.0206 mol/L. What is the average rate of consumption in moles per litre per second of reactant B?

#### What Is Required?

You need to determine the average rate of consumption of reactant B over the first 14.7 minutes of the reaction.

#### What Is Given?

You know the initial concentration of A:  $[A]_{\text{initial}} = 0.0415 \text{ mol/L}$

You know the final concentration of A:  $[A]_{\text{final}} = 0.0206 \text{ mol/L}$

You know the balanced chemical equation for the reaction between A and B.

Plan Your Strategy	Act on Your Strategy
Convert the units for time elapsed, $\Delta t$ , from minutes to seconds.	$\Delta t = 14.7 \text{ min}$ $= 14.7 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}}$ $= 882 \text{ s}$
Use the formula $\text{average rate} = \frac{\Delta[A]}{\Delta t}$ to determine the rate of consumption of reactant A.	$\frac{\Delta A}{\Delta t} = \frac{0.0206 \text{ mol/L} - 0.0415 \text{ mol/L}}{882 \text{ s}}$ $= -2.3696 \times 10^{-5} \text{ mol/L} \cdot \text{s}$
The rate of consumption of reactants is in proportion to the mole ratio given in the balanced chemical equation. Use the mole ratio in the balanced equation A:B to determine the rate of consumption of reactant B.	$n_B : n_A = 2 : 1$ $\frac{\Delta[B]}{\Delta t} = 2 \left( \frac{\Delta[A]}{\Delta t} \right)$ $= 2(-2.3696 \times 10^{-5} \text{ mol/L} \cdot \text{s})$ $= -4.7392 \times 10^{-5} \text{ mol/L} \cdot \text{s}$ $= -4.74 \times 10^{-5} \text{ mol/L} \cdot \text{s}$  The average rate of reaction of B is $-4.74 \times 10^{-5} \text{ mol/L} \cdot \text{s}$ . The value is negative because the concentration is decreasing. The rate of consumption is the opposite of the rate at which it is decreasing so the rate of consumption is $+4.74 \times 10^{-5} \text{ mol/L} \cdot \text{s}$ .

### Check Your Solution

The answer is correctly expressed in the proper units and to the correct number of significant digits. The rate of consumption of reactant B is twice that of reactant A as shown by the mol ratio in the balanced equation.

2. Calculate the average rate of a reaction, given the following data:

### Reaction Data

Time (s)	Concentration (mol/L)
60	$5.00 \times 10^{-2}$
85	$3.25 \times 10^{-2}$

### What Is Required?

You need to calculate the average rate of reaction from the given data.

### What Is Given?

You know the concentration at two different times.

You know the initial concentration:  $c_{\text{initial}} = 5.00 \times 10^{-2}$  mol/L

You know the final concentration:  $c_{\text{final}} = 3.25 \times 10^{-2}$  mol/L

You know the initial time:  $t_{\text{initial}} = 60$  s

You know the final time:  $t_{\text{final}} = 85$  s

Plan Your Strategy	Act on Your Strategy
Determine the time elapsed, $\Delta t$ .	$\begin{aligned}\Delta t &= t_{\text{final}} - t_{\text{initial}} \\ &= 85 \text{ s} - 60 \text{ s} \\ &= 25 \text{ s}\end{aligned}$
Use the formula $\text{rate} = \frac{\Delta \text{concentration}}{\Delta t}$ to calculate the average rate of reaction.	$\begin{aligned}\text{rate} &= \frac{\Delta \text{concentration}}{\Delta t} \\ &= \frac{c_{\text{final}} - c_{\text{initial}}}{\Delta t} \\ &= \frac{3.25 \times 10^{-2} \text{ mol/L} - 5.00 \times 10^{-2} \text{ mol/L}}{25 \text{ s}} \\ &= -7.0 \times 10^{-4} \text{ mol/L} \cdot \text{s}\end{aligned}$ <p>The average rate of the reaction is <math>-7.0 \times 10^{-4}</math> mol/L•s</p>

### Check Your Solution

The reaction rate is negative since the concentration is decreasing. The answer has the correct number of significant digits and seems reasonable.

3. For the reaction shown below, the average rate of formation of  $\text{CO}_2(\text{g})$  is  $5.50 \times 10^{-4}$  mol/s.



- a. What amount in moles of  $\text{CO}_2(\text{g})$  is formed in 5.00 min?

### What Is Required?

You need to calculate the amount in moles of  $\text{CO}_2(\text{g})$  produced in 5.00 min.

### What Is Given?

You know the balanced chemical equation for the reaction.

You know the average rate of formation of  $\text{CO}_2(\text{g})$ : rate =  $5.50 \times 10^{-4}$  mol/s

You know the time elapsed:  $t = 5.00$  min

Plan Your Strategy	Act on Your Strategy
Convert the units for time elapsed, $t$ , from minutes to seconds.	$t = 5.00 \text{ min}$ $= 5.00 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}}$ $= 300 \text{ s}$
Use the expression $n = \frac{\Delta \text{concentration}}{\Delta t} \times t$ to calculate the amount in moles, $n$ , of $\text{CO}_2(\text{g})$ .	$n_{\text{CO}_2} = \frac{\Delta \text{concentration}}{\Delta t} \times t$ $= \frac{5.50 \times 10^{-4} \text{ mol}}{1 \cancel{\text{s}}} \times 300 \cancel{\text{s}}$ $= 0.165 \text{ mol}$

- b. How does this compare with the amount of  $\text{Br}_2(\text{aq})$  that reacts in the same time?

### What Is Required?

You need to compare the amount of  $\text{CO}_2(\text{g})$  produced to the amount of  $\text{Br}_2(\text{aq})$  consumed-

### What Is Given?

You know the amount of  $\text{CO}_2(\text{g})$  produced.

You know the balanced chemical equation.

Plan Your Strategy	Act on Your Strategy
The rate of consumption of reactants is in proportion to the mole ratio given in the balanced equation. Determine the mole ratio from the balanced equation.	$\frac{n_{\text{Br}_2}}{n_{\text{CO}_2}} = \frac{1}{1}$



Determine the rate at which Br <sub>2</sub> (aq) reacts.	$\frac{\Delta[\text{Br}_2]}{\Delta t} = \frac{\Delta[\text{CO}_2]}{\Delta t}$ <p>Br<sub>2</sub>(aq) will be consumed at the same rate as CO<sub>2</sub> is produced. Therefore, 0.165 mol of Br<sub>2</sub>(aq) reacts in this same time interval.</p>
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### Check Your Solution

The answer has the correct number of significant digits. The equation shows a 1:1 mol ratio between Br<sub>2</sub>(aq) and CO<sub>2</sub>(g). The answer is reasonable.

4. The concentration of a reactant is  $4.0 \times 10^{-2}$  mol/L at  $t = 2.0$  min. If the average rate of consumption of the reactant from  $t = 1.5$  min to  $t = 2.0$  min was 0.045 mol/L•s, what was the concentration of this reactant at  $t = 1.5$  min?

### What Is Required?

You need to determine the concentration of a reactant after 1.5 minutes have elapsed in the reaction.

### What Is Given?

You know the average rate of reaction: rate = 0.045 mol/L•s

You know the final concentration:  $c_{\text{final}} = 4.0 \times 10^{-2}$  mol/L

You know the initial time:  $t_{\text{initial}} = 1.5$  min

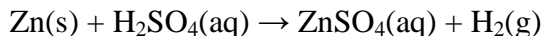
You know the final time:  $t_{\text{final}} = 2.0$  min

Plan Your Strategy	Act on Your Strategy
Convert the units for $t_{\text{initial}}$ and $t_{\text{final}}$ from minutes to seconds.	$t_{\text{initial}} = 1.5 \text{ min}$ $= 1.5 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}}$ $= 90 \text{ s}$  $t_{\text{final}} = 2.0 \text{ min}$ $= 2.0 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}}$ $= 120 \text{ s}$
Determine the time elapsed, $\Delta t$ .	$\Delta t = t_{\text{final}} - t_{\text{initial}}$ $= 120 \text{ s} - 90 \text{ s}$ $= 30 \text{ s}$
Use the formula $\text{average rate} = -\frac{\Delta \text{concentration}}{\Delta t}$ to calculate the initial concentration, $c_{\text{initial}}$ . The rate is negative because the amount of reactant is decreasing.	$\text{average rate} = -\frac{\Delta \text{concentration}}{\Delta t}$ $\text{average rate} = -\frac{(c_{\text{final}} - c_{\text{initial}})}{\Delta t}$ $0.045 \text{ mol/L} \cdot \text{s} = \frac{-(4.0 \times 10^{-2} \text{ mol/L} - c_{\text{initial}})}{30 \text{ s}}$ $c_{\text{initial}} = 1.346 \text{ mol/L}$ $= 1.4 \text{ mol/L}$  The initial concentration of the reactant was 1.4 mol/L.

### Check Your Solution

The initial concentration is greater than the final concentration. The answer has the correct number of significant digits and seems reasonable.

5. A zinc electrode is immersed in dilute sulfuric acid at 35.0°C and the following reaction occurs:



The volume of  $\text{H}_2(\text{g})$  present at  $t = 1.0$  min is 30.0 mL and at  $t = 1.4$  min is 42.0 mL. What is the average rate of formation of  $\text{H}_2(\text{g})$  over this period of time measured in litres per second?

### What Is Required?

You must determine the average rate of formation of the product,  $\text{H}_2(\text{g})$ .

### What Is Given?

You know the initial volume:  $V_{\text{initial}} = 30.0$  mL

You know the final volume:  $V_{\text{final}} = 42.0$  mL

You know the initial time:  $t_{\text{initial}} = 1.0$  min

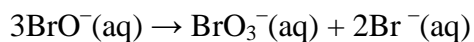
You know the final time:  $t_{\text{final}} = 1.4$  min

Plan Your Strategy	Act on Your Strategy
Calculate $\Delta t$ and convert the units from minutes to seconds.	$\begin{aligned}\Delta t &= t_{\text{final}} - t_{\text{initial}} \\ &= 1.4 \text{ min} - 1.0 \text{ min} \\ &= 0.4 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}} \\ &= 24 \text{ s}\end{aligned}$
Use the formula $\text{rate} = \frac{\Delta \text{quantity}}{\Delta t}$ to determine the average rate of formation of $\text{H}_2(\text{g})$ .	$\begin{aligned}\text{rate of formation} &= \frac{\Delta \text{quantity}}{\Delta t} \\ &= \frac{\Delta V}{\Delta t} \\ &= \frac{V_{\text{final}} - V_{\text{initial}}}{\Delta t} \\ &= \frac{(42.0 \text{ mL} - 30.0 \text{ mL})}{24 \text{ s}} \\ &= \frac{12 \text{ mL}}{24 \text{ s}} \\ &= 0.50 \text{ mL/s} \\ &= 5.0 \times 10^{-4} \text{ L/s}\end{aligned}$

### Check Your Solution

The rate is positive since  $\text{H}_2(\text{g})$  is produced. The correct number of significant digits is shown.

6. For the reaction shown below, the instantaneous rate of formation of  $\text{Br}^-(\text{aq})$  is  $0.12 \text{ mol/L}\cdot\text{s}$  at  $t = 2.0 \text{ min}$ .



What are the instantaneous rates of formation of  $\text{BrO}_3^-(\text{aq})$  and consumption of  $\text{BrO}^-(\text{aq})$ ?

### What Is Required?

You need to calculate the instantaneous rate formation of  $\text{BrO}_3^-(\text{aq})$  and rate of consumption of  $\text{BrO}^-(\text{aq})$ ?

### What Is Given?

You know the balanced chemical equation for the reaction.

You know the instantaneous rate of formation of  $\text{Br}^-(\text{aq})$  is  $0.12 \text{ mol/L}\cdot\text{s}$ .

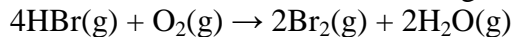
Plan Your Strategy	Act on Your Strategy
The rate of formation of products and the rate of consumption of reactants is in proportion to the mole ratios given in the balanced equation. Determine the mole ratios $\text{BrO}_3^-:\text{Br}^-$ and $\text{BrO}^-:\text{Br}^-$ from the balanced equation.	$\frac{\text{rate of production of } \text{BrO}_3^-(\text{aq})}{\text{rate of production of } \text{Br}^-(\text{aq})} = \frac{n_{\text{BrO}_3^-}}{n_{\text{Br}^-}}$ $\frac{-\text{rate of disappearance of } \text{BrO}^-(\text{aq})}{\text{rate of production of } \text{Br}^-(\text{aq})} = \frac{n_{\text{BrO}^-}}{n_{\text{Br}^-}}$
Use the first ratio to calculate the instantaneous rate of formation of $\text{BrO}_3^-(\text{aq})$ .	<p>For every 2 mol of <math>\text{Br}^-(\text{aq})</math> produced, 1 mol of <math>\text{BrO}_3^-(\text{aq})</math> is produced.</p> $\frac{\text{rate of production of } \text{BrO}_3^-(\text{aq})}{0.12 \text{ mol/L}\cdot\text{s}} = \frac{1 \text{ mol } \text{BrO}_3^-(\text{aq})}{2 \text{ mol } \text{Br}^-(\text{aq})}$ $\text{rate of production of } \text{BrO}_3^-(\text{aq}) = \frac{1}{2} \times 0.12 \text{ mol/L}\cdot\text{s}$ $= 0.06 \text{ mol/L}\cdot\text{s}$ <p><math>\text{BrO}_3^-(\text{aq})</math> is produced half as fast as <math>\text{Br}^-(\text{aq})</math>.</p>

Use the second ratio to calculate the instantaneous rate of consumption of $\text{BrO}^-$ (aq).	<p>For every 2 mol <math>\text{Br}^-</math> (aq) that is produced, 3 mol <math>\text{BrO}^-</math> (aq) is consumed.</p> $\frac{-\text{rate of disappearance of } \text{BrO}^- \text{ (aq)}}{0.12 \text{ mol/L} \cdot \text{s}} = \frac{3 \text{ mol } \text{BrO}^- \text{ (aq)}}{2 \text{ mol } \text{Br}^- \text{ (aq)}}$ $\text{rate of disappearance of } \text{BrO}^- \text{ (aq)} = \frac{3}{2} \times 0.12 \text{ mol/L} \cdot \text{s}$ $= 0.18 \text{ mol/L} \cdot \text{s}$ <p><math>\text{BrO}^-</math> (aq) disappears 1.5 times faster than <math>\text{Br}^-</math> (aq) is produced.</p>
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### Check Your Solution

The answers are reasonable based upon the balanced equation. The correct number of significant digits is shown.

7. The data in the table below were obtained for the following reaction:



### Reaction Data

Time (s)	[HBr] (mol/L)	[Br <sub>2</sub> ] (mol/L)
0.00	0.42	0.00
50.0	0.26	?

- What is the average rate of consumption of HBr(g) over 50.0 s?
- What is the molar concentration of Br<sub>2</sub>(g) at  $t = 50.0$  s?

### What Is Required?

You need to calculate rate of consumption of HBr(g) and the final molar concentration of Br<sub>2</sub>(g).

### What Is Given?

You know the balanced chemical equation for the reaction.

You know the initial concentration of HBr(g):  $c_{\text{initial}} = 0.42$  mol/L

You know the final concentration of HBr(g):  $c_{\text{final}} = 0.26$  mol/L

You know the initial concentration of Br<sub>2</sub>(g):  $c_{\text{initial}} = 0.00$  mol/L

You know the initial time:  $t_{\text{initial}} = 0.00$  s

You know the final time:  $t_{\text{final}} = 50.0$  s

Plan Your Strategy	Act on Your Strategy
<b>a.</b> Use the formula: $\text{rate} = \frac{\Delta \text{concentration}}{\Delta t}$ to determine the rate of consumption of HBr(g).	<b>a.</b> rate of consumption of HBr(g): $\begin{aligned} \text{rate} &= \frac{\Delta \text{concentration}}{\Delta t} \\ &= \frac{c_{\text{final}} - c_{\text{initial}}}{\Delta t} \\ &= \frac{0.26 \text{ mol/L} - 0.42 \text{ mol/L}}{50 \text{ s}} \\ &= -0.0032 \text{ mol/L} \cdot \text{s} \end{aligned}$  The average rate of consumption of HBr(g) over 50.0 s is 0.0032 mol/L•s.
<b>b.</b> The rate of formation of products is in proportion to the mole ratio given in the balanced chemical equation. Determine the mole ratio HBr:Br <sub>2</sub> from the balanced equation.	<b>b.</b> $\frac{n_{\text{HBr}}}{n_{\text{Br}_2}} = \frac{4}{2}$

Use the mole ratio to calculate the rate of formation of Br <sub>2</sub> .	<p>For every 4 mol HBr(g) consumed, 2 mol Br<sub>2</sub>(g) is produced.</p> $\frac{-\Delta[\text{HBr}]}{\Delta[\text{Br}_2]} = \frac{4}{2}$ $\Delta[\text{Br}_2] = \left(\frac{2}{4}\right) (-\Delta[\text{HBr}])$ $= (0.5) [-(-0.0032 \text{ mol/L} \cdot \text{s})]$ $= 0.0016 \text{ mol/L} \cdot \text{s}$
Use the formula: $\text{rate} = \frac{\Delta \text{concentration}}{\Delta t}$ to determine the final concentration, $c_{\text{final}}$ , of Br <sub>2</sub> (g) at 50 s.	$\text{rate} = \frac{\Delta \text{concentration}}{\Delta t}$ $\text{rate} = \frac{(c_{\text{final}} - c_{\text{initial}})}{\Delta t}$ $0.0016 \text{ mol/L} \cdot \text{s} = \frac{(c_{\text{final}} - 0.00 \text{ mol/L})}{50 \text{ s}}$ $c_{\text{final}} = 0.080 \text{ mol/L}$ <p>The final concentration of Br<sub>2</sub>(g) is 0.080 mol/L.</p>

### Check Your Solution

The average rate of reaction has the correct number of significant digits and seems reasonable. The final concentration of Br<sub>2</sub>(g) is greater than the initial concentration.

8. Consider the following reaction:  $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g})$

Over a period of 1.80 min, the average rate of formation of  $\text{NO}(\text{g})$  is  $1.04 \text{ mol/L}\cdot\text{s}$ .  
What was the amount in moles of  $\text{O}_2(\text{g})$  consumed over this period of time?

### What Is Required?

You need to determine the amount in moles of  $\text{O}_2(\text{g})$  consumed.

### What Is Given?

You know the balanced chemical equation for the reaction.

Plan Your Strategy	Act on Your Strategy
Convert the unit of time from minutes to seconds.	$\Delta t = 1.8 \text{ min}$ $= 1.8 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}}$ $= 108 \text{ s}$
The rate of formation of products is in proportion to the mole ratio given in the balanced chemical equation. Determine the mole ratio $\text{O}_2(\text{g})\text{:NO}(\text{g})$	$\frac{n_{\text{O}_2}}{n_{\text{NO}_2}} = \frac{5}{4}$
Calculate the rate of consumption of $\text{O}_2(\text{g})$ .	<p>For every 4 mol <math>\text{NO}(\text{g})</math> produced, 5 mol <math>\text{O}_2(\text{g})</math> is consumed.</p> $\frac{\text{rate of consumption of } \text{O}_2(\text{g})}{\text{rate of production of } \text{NO}_2(\text{g})} = \frac{5 \text{ mol } \text{O}_2(\text{g})}{4 \text{ mol } \text{NO}_2(\text{g})}$ $\frac{\text{rate of consumption of } \text{O}_2(\text{g})}{1.04 \text{ mol/L}\cdot\text{s}} = \frac{5 \text{ mol } \text{O}_2(\text{g})}{4 \text{ mol } \text{NO}_2(\text{g})}$ $\text{rate of consumption of } \text{O}_2(\text{g}) = 1.04 \text{ mol/L}\cdot\text{s} \times 1.25$ $= 1.30 \text{ mol/L}\cdot\text{s}$ <p>The rate of consumption of <math>\text{O}_2</math> is <math>1.30 \text{ mol/L}\cdot\text{s}</math>.</p>
Calculate the amount of $\text{O}_2(\text{g})$ consumed after 108 s.	$\Delta \text{O}_2 = 1.30 \text{ mol/L}\cdot\cancel{\text{s}} \times 108 \cancel{\text{s}}$ $= 140.4 \text{ mol/L}$ $= 1.40 \times 10^2 \text{ mol/L}$ <p><math>1.40 \times 10^2 \text{ mol}</math> of <math>\text{O}_2(\text{g})</math> was consumed.</p>

### Check Your Solution

The concentration of the reactant  $\text{O}_2(\text{g})$  decreased in proportion to the mole ratio in the balanced equation. The answer is reasonable, and the correct number of significant digits is shown.



9. A mass of 0.50 g of sodium metal reacts with water in 90.0 s.



- Express the rate of consumption of Na(s) in moles per second.
- Calculate the rate at which H<sub>2</sub>(g) is generated, in litres per second, at 30.0°C and 102.4 kPa. Assume that hydrogen is an ideal gas.

### What Is Required?

- You need to determine the rate of consumption of Na(s) at a given temperature and pressure.
- You need to determine the rate of formation of H<sub>2</sub>(g) at a given temperature and pressure.

### What Is Given?

You know the balanced chemical equation for the reaction.

You know the mass of Na(s):  $m = 0.50 \text{ g}$

You know the temperature:  $T = 30.0^\circ\text{C}$

You know the pressure:  $P = 102.4 \text{ kPa}$

You know the elapsed time:  $\Delta t = 90.0 \text{ s}$

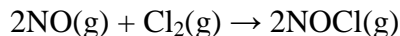
Plan Your Strategy	Act on Your Strategy
<p>a. Using the mass and molar mass of sodium, use the formula <math>n = \frac{m}{M}</math> to determine the amount in moles, <math>n</math>, of Na(s).</p>	$n = \frac{m}{M}$ $n_{\text{Na}} = \frac{0.50 \cancel{\text{g}}}{22.99 \cancel{\text{g}}/\text{mol}}$ $= 0.021\,75 \text{ mol}$
<p>Use the formula <math>\text{rate} = \frac{\Delta \text{quantity}}{\Delta t}</math> to calculate the rate at which Na(s) is consumed in the reaction.</p>	$\text{rate of consumption of Na(s)} = \frac{\Delta \text{quantity}}{\Delta t}$ $= \frac{0.021\,75 \text{ mol}}{90.0 \text{ s}}$ $= 2.42 \times 10^{-4} \text{ mol/s}$ <p>Na(s) is consumed at the rate of <math>2.42 \times 10^{-4} \text{ mol/s}</math>.</p>
<p>The rate of formation of products and the rate of consumption of reactants is in proportion to the mol ratio given in the balanced equation. Determine the mole ratio of Na(s):H<sub>2</sub>(g) from the balanced equation.</p>	$\frac{n_{\text{Na}}}{n_{\text{H}_2}} = \frac{2}{1}$

Calculate the rate of formation of $\text{H}_2(\text{g})$ in mol/s.	<p>For every 1 mol <math>\text{H}_2(\text{g})</math> produced, 2 mol <math>\text{Na}(\text{s})</math> is consumed.</p> $\frac{\text{rate of formation of } \text{H}_2(\text{g})}{\text{rate of consumption of } \text{Na}(\text{s})} = \frac{n_{\text{H}_2}}{n_{\text{Na}}}$ $\frac{\text{rate of formation of } \text{H}_2(\text{g})}{2.42 \times 10^{-4} \text{ mol/s}} = \frac{1}{2}$ $\text{rate of formation of } \text{H}_2(\text{g}) = \frac{2.42 \times 10^{-4} \text{ mol/s}}{2}$ $= 1.21 \times 10^{-4} \text{ mol/s}$
Use the ideal gas equation, $PV = nRT$ , to convert the amount in moles of $\text{H}_2(\text{g})$ to volume of $\text{H}_2(\text{g})$ . Express the rate of formation of $\text{H}_2(\text{g})$ in L/s.	$V = \frac{nRT}{P}$ $= \frac{(1.21 \times 10^{-4} \text{ mol}) \left( 8.314 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}} \right) (303.15 \text{ K})}{102.4 \text{ kPa}}$ $= 2.98 \times 10^{-3} \text{ L}$ <p>The rate of formation of <math>\text{H}_2(\text{g})</math> is <math>2.98 \times 10^{-3} \text{ L/s}</math>.</p>

### Check Your Solution

The rate of formation of  $\text{H}_2(\text{g})$  is positive. The number of significant digits is shown correctly and the answer seems reasonable.

10. For the reaction shown below, the instantaneous change in concentration of NO(g) is 1.4 mol/L•s.



What is the rate at which Cl<sub>2</sub>(g) is consumed and the rate at which NOCl(g) is formed at this time?

### What Is Required?

You need to determine the rate at which Cl<sub>2</sub>(g) is consumed

You need to determine the rate at which NOCl(g) is produced.

### What Is Given?

You know the balanced chemical equation.

The instantaneous change in concentration of NO(g) is –1.4 mol/L•s.

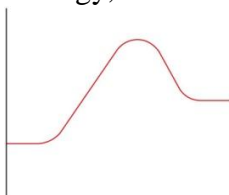
Plan Your Strategy	Act on Your Strategy
The rate of formation of products and the rate of consumption of reactants is in proportion to the mole ratios given in the balanced chemical equation. Determine the mole ratios Cl <sub>2</sub> :NO and NOCl:NO from the balanced equation.	$\frac{n_{\text{Cl}_2}}{n_{\text{NO}}} = \frac{1}{2} \qquad \frac{n_{\text{NOCl}}}{n_{\text{NO}}} = \frac{2}{2}$
Calculate the rate of consumption of Cl <sub>2</sub> (g).	<p>For every 1 mol Cl<sub>2</sub>(g) consumed, 2 mol NOCl(g) is produced.</p> $\frac{\text{rate of consumption Cl}_2\text{(g)}}{\text{rate of consumption NO(g)}} = \frac{n_{\text{Cl}_2}}{n_{\text{NO}}}$ $\frac{\text{rate of consumption Cl}_2\text{(g)}}{-1.4 \text{ mol/L} \cdot \text{s}} = \frac{1}{2}$ $\text{rate of consumption Cl}_2\text{(g)} = \frac{-1.4 \text{ mol/L} \cdot \text{s}}{2}$ $= -0.7 \text{ mol/L} \cdot \text{s}$ <p>Cl<sub>2</sub>(g) is consumed at a rate of 0.70 mol/L•s.</p>
Calculate the rate of formation of NOCl(g).	<p>Since NO(g) and NOCl(g) have a molar ratio 2:2, NOCl(g) is produced at the same rate that NO(g) is consumed. NOCl(g) is therefore produced at a rate of 1.4 mol/L•s.</p>

### Check Your Solution

The rates of consumption and formation agree with the mol ratio in the balanced equation. The answers are expressed to the correct number of significant digits.

## Representing a Reaction with a Potential Energy Diagram (Student textbook page 371)

11. Complete the following potential energy diagram by adding the following labels: an appropriate label for the  $x$ -axis and  $y$ -axis,  $E_{a(\text{fwd})}$ ,  $E_{a(\text{rev})}$ ,  $\Delta H_r$ .
- Is the forward reaction endothermic or exothermic?
  - Which has the higher potential energy, the reactants or the products?



### What Is Required?

You need to label the given potential energy diagram, determine if the reaction is exothermic or endothermic and indicate which has the higher potential energy, reactants or products.

### What Is Given?

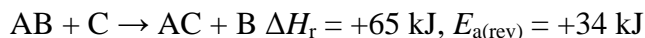
You have an unlabelled potential energy diagram.

Plan Your Strategy	Act on Your Strategy
<p>The <math>y</math>-axis is an energy scale, and the <math>x</math>-axis follows the reaction progress. The reactants will be shown as the first level portion, and the products will be the second level portion of the diagram. The activation energy in the forward direction, <math>E_{a(\text{fwd})}</math>, is the difference between the reactant energy and the transition state at the peak of the diagram. The activation energy for the reverse reaction, <math>E_{a(\text{rev})}</math>, is the difference between the product energy and transition state at the peak of the diagram. <math>\Delta H_r</math> is the difference between the potential energy of the reactant and the potential energy of the product.</p>	<p><b>a.</b> The reaction is endothermic.  <b>b.</b> The products have a higher potential energy than the reactants.</p>

### Check Your Solution

The potential energy of the reactant is lower than the potential energy of the product. This corresponds to an endothermic reaction.

12. Consider the following reaction:



Draw and label a potential energy diagram for this reaction. Calculate and label  $E_{\text{a}(\text{fwd})}$ .

### What Is Required?

You need to draw a potential energy diagram labelling the  $x$ -axis and  $y$ -axis, the transition state, and  $E_{\text{a}(\text{fwd})}$ .

### What Is Given?

You know the balanced chemical equation for the reaction.

You know the activation energy of the reverse reaction:  $E_{\text{a}(\text{rev})} = +34 \text{ kJ}$

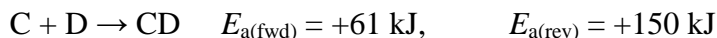
You know the enthalpy change of the forward reaction  $\Delta H_{\text{r}} = +65 \text{ kJ}$

Plan Your Strategy	Act on Your Strategy
Use the formula $\Delta H_{\text{r}} = E_{\text{a}(\text{fwd})} - E_{\text{a}(\text{rev})}$ to calculate $E_{\text{a}(\text{fwd})}$ .	$\Delta H_{\text{r}} = E_{\text{a}(\text{fwd})} - E_{\text{a}(\text{rev})}$ $+65 \text{ kJ} = E_{\text{a}(\text{fwd})} - (+34 \text{ kJ})$ $E_{\text{a}(\text{fwd})} = +99 \text{ kJ}$
Draw and label the potential energy diagram.	

### Check Your Solution

Since the sign of  $\Delta H_{\text{r}}$  is positive, the reaction is endothermic and the potential energy diagram shows the reactants at a lower energy than the products.

13. Consider the reaction below:



Draw and label a potential energy diagram for this reaction. Calculate and label  $\Delta H_{\text{r}}$ .

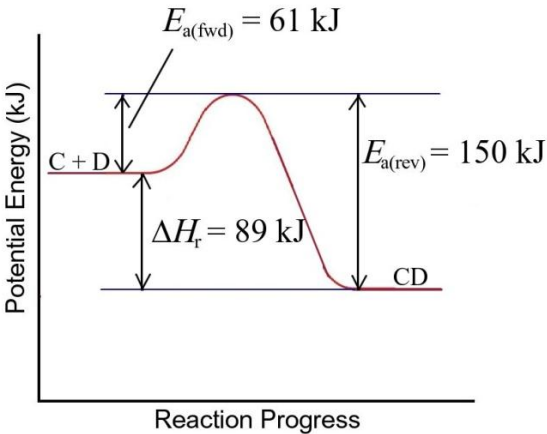
### What Is Required?

You need to draw and label a potential energy diagram for this reaction and calculate and label  $\Delta H_{\text{r}}$ .

### What Is Given?

You know the activation energy of the forward reaction:  $E_{\text{a(fwd)}} = +61 \text{ kJ}$

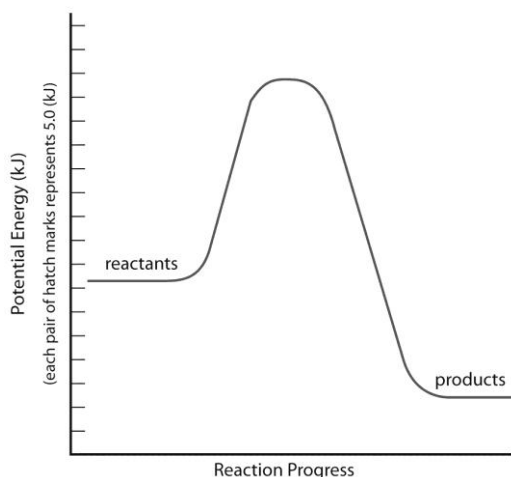
You know the activation energy of the reverse reaction:  $E_{\text{a(rev)}} = +150 \text{ kJ}$

Plan Your Strategy	Act on Your Strategy
Use the formula $\Delta H_{\text{r}} = E_{\text{a(fwd)}} - E_{\text{a(rev)}}$ to calculate $\Delta H_{\text{r}}$ .	$\Delta H_{\text{r}} = E_{\text{a(fwd)}} - E_{\text{a(rev)}}$ $= +61 \text{ kJ} - (+150 \text{ kJ})$ $= -89 \text{ kJ}$
Draw and label the potential energy diagram.	 <p><math>E_{\text{a(rev)}}</math> is the energy difference between the transition state and the products (+150 kJ). <math>\Delta H_{\text{r}}</math> is the enthalpy of reaction between the reactants and the products (-89 kJ).</p>

### Check Your Solution

Since the sign of  $\Delta H_{\text{r}}$  is negative, the reaction is exothermic. The diagram correctly shows the potential energy of the product lower than the potential energy of the reactant.

14. Using the potential energy diagram below, estimate the values for  $E_{a(\text{fwd})}$ ,  $E_{a(\text{rev})}$ , and  $\Delta H_r$ . Is the reaction endothermic or exothermic?



### What Is Required?

You need to estimate the values for  $E_{a(\text{fwd})}$ ,  $E_{a(\text{rev})}$ , and  $\Delta H_r$  based upon the given potential energy diagram and indicate if the reaction endothermic or exothermic.

### What Is Given?

You have a potential energy diagram.

Plan Your Strategy	Act on Your Strategy
Refer to the scale on the y-axis and determine the difference in energy between the reactants and the transition state, $E_{a(\text{fwd})}$ ; the difference in energy between the products and the transition state, $E_{a(\text{rev})}$ ; and the difference between the energy of the reactants and products, $\Delta H_r$ .	<p>Estimated values for <math>E_{a(\text{fwd})}</math>, <math>E_{a(\text{rev})}</math>, and <math>\Delta H_r</math> are:</p> <p><math>E_{a(\text{fwd})} = +42 \text{ kJ}</math>  <math>E_{a(\text{rev})} = +67 \text{ kJ}</math>  <math>\Delta H_r = -25 \text{ kJ}</math></p> <p>The reaction is exothermic.</p>

### Check Your Solution

Check to be certain that you have counted the correct number of divisions when determining the energies on the y-axis.

15. In the upper atmosphere, oxygen exists as  $\text{O}_2(\text{g})$ , as ozone,  $\text{O}_3(\text{g})$ , and as individual oxygen atoms,  $\text{O}(\text{g})$ . Ozone and atomic oxygen react to form two molecules of oxygen gas. The enthalpy change is  $-392 \text{ kJ}$  and the activation energy is  $+19.0 \text{ kJ}$ . Draw and label a potential energy diagram. Include a value for  $E_{\text{a}(\text{rev})}$ .

### What Is Required?

You need to determine the value of  $E_{\text{a}(\text{rev})}$  for a reaction and draw and label a potential energy diagram.

### What Is Given?

You know the activation energy of the forward reaction:  $E_{\text{a}(\text{fwd})} = +19 \text{ kJ}$

You know the enthalpy change of the forward reaction:  $\Delta H_{\text{r}} = -392 \text{ kJ}$

Plan Your Strategy	Act on Your Strategy
Use the formula $\Delta H_{\text{r}} = E_{\text{a}(\text{fwd})} - E_{\text{a}(\text{rev})}$ to determine $E_{\text{a}(\text{rev})}$ .	$\Delta H_{\text{r}} = E_{\text{a}(\text{fwd})} - E_{\text{a}(\text{rev})}$ $-392 \text{ kJ} = +19 \text{ kJ} - E_{\text{a}(\text{rev})}$ $E_{\text{a}(\text{rev})} = +411 \text{ kJ}$
Draw and label a potential energy diagram.	

### Check Your Solution

Since the sign of  $\Delta H_{\text{r}}$  is negative, the reaction is exothermic. The diagram correctly shows the potential energy of the product lower than the potential energy of the reactant.



16. For a reaction, on an arbitrary scale, the potential energies are as follows: activated complex, +112 kJ; reactants, +36 kJ; products, +78 kJ.
- Determine the activation energy and the enthalpy change for the reaction.
  - Draw a labelled potential energy diagram for the reaction, indicating the relative energies of the reactants, products, and activated complex.

### What Is Required?

You need to determine the activation energy,  $E_a$ , and the enthalpy change,  $\Delta H_r$ , for the reaction.

You need to draw a potential energy diagram for the reaction.

### What Is Given?

You know the potential energies: activated complex, +112 kJ; reactants, +36 kJ; products, +78 kJ.

Plan Your Strategy	Act on Your Strategy
<p>a. The activation energy is the difference between the potential energy of the activated complex and the potential energy of the reactants. The enthalpy change is the difference in potential energy between the products and reactants. Estimate the energy scale (y-axis) that will include the calculated differences in energy.</p>	$E_a = 112 \text{ kJ} - 36 \text{ kJ}$ $= +76 \text{ kJ}$ $\Delta H = 78 \text{ kJ} - 36 \text{ kJ}$ $= +42 \text{ kJ}$
<p>b. Draw and label the potential energy diagram.</p>	

### Check Your Solution

The sign of the enthalpy change is positive, and the diagram will show the potential energy of the reactants lower than the potential energy of the products.

17. Refer to the list of molar enthalpies of combustion for hydrocarbons in **Table 5.4** (student textbook page 295).
- Write the balanced thermochemical equation for the combustion of methane gas,  $\text{CH}_4(\text{g})$ .
  - Draw a potential energy diagram that would reasonably represent this combustion reaction. Indicate the  $\Delta H_{\text{comb}}$  and a molecular structure that could represent an activated complex in this potential energy diagram.

### What Is Required?

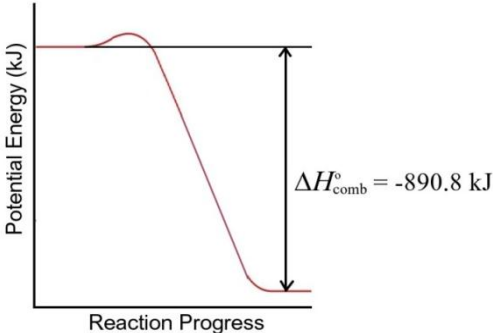
You need to write the thermochemical equation for the combustion of methane and draw a reasonable energy profile diagram that could represent the process. Include a molecular structure that could represent an activated complex in the potential energy diagram.

### What Is Given?

From **Table 5.4**, you know the molar enthalpy of combustion for methane:

$$\Delta H_{\text{comb}}^{\circ} = -890.8 \text{ kJ/mol}$$

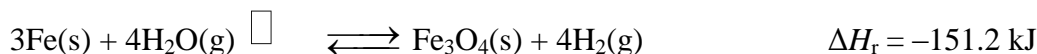
The products for complete combustion are  $\text{H}_2\text{O}(\ell)$  and  $\text{CO}_2(\text{g})$ .

Plan Your Strategy	Act on Your Strategy
a. <b>Table 5.4</b> gives a negative value for the heat of combustion so the equation for the combustion reaction shows the heat term as a product.	$\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\ell) + 890.8 \text{ kJ}$
b. Draw and label potential energy diagram for the reaction including a molecular structure that could represent an activated complex.	 <p>The activated complex would show an unstable association of one <math>\text{CH}_4(\text{g})</math> molecule and <math>\text{O}_2(\text{g})</math> molecule with partial bonds.</p>

### Check Your Solution

The potential energy diagram should match the given information. An example of the burning of a hydrocarbon is given in the student text page 367.

18. When steam is passed over hot iron, a reaction occurs as shown below.



The activation energy for the reverse reaction,  $E_{a(\text{rev})}$ , is +200.71 kJ.

- Calculate the activation energy for the forward reaction.
- Draw a labelled potential energy diagram showing the enthalpy change, and the activation energies for the forward and reverse reactions.

### What Is Required?

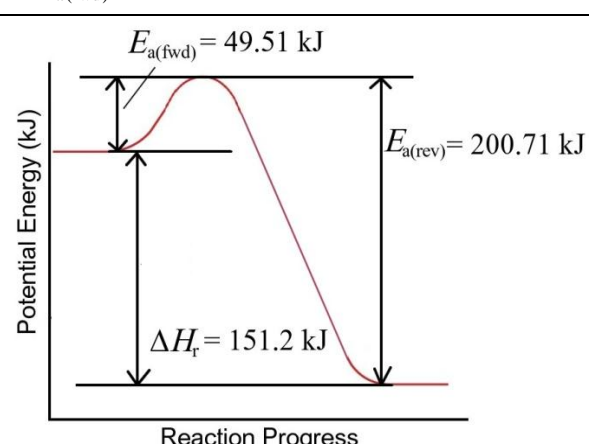
You need to calculate the activation energy for a forward reaction.

You need to draw a labelled potential energy diagram showing the enthalpy change, and the activation energies for the forward and reverse reactions.

### What Is Given?

You know the enthalpy change of the forward reaction:  $\Delta H_r = -151.2 \text{ kJ/mol}$

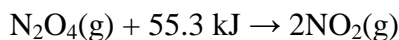
You know the activation energy of the reverse reaction:  $E_{a(\text{rev})} = +200.71 \text{ kJ}$

Plan Your Strategy	Act on Your Strategy
a. Use the formula $\Delta H_r = E_{a(\text{fwd})} - E_{a(\text{rev})}$ to calculate $E_{a(\text{fwd})}$ .	$\Delta H_r = E_{a(\text{fwd})} - E_{a(\text{rev})}$ $-151.2 \text{ kJ} = E_{a(\text{fwd})} - (+200.71 \text{ kJ})$ $E_{a(\text{fwd})} = +49.5 \text{ kJ}$
b. Draw a labelled energy profile diagram.	 <p>Diagram shows that the reactants have higher potential energy than the products. The activated complex, having the highest potential energy, is positioned between the reactants and products.</p>

### Check Your Solution

Check that the labels are all included. The reaction is exothermic so the potential energy of the products should be lower than the potential energy of the reactants.

19. The decomposition of dinitrogen tetroxide(g),  $\text{N}_2\text{O}_4(\text{g})$ , to nitrogen dioxide,  $\text{NO}_2(\text{g})$ , is a reversible reaction. The activation energy for the decomposition reaction is +58.6 kJ.



Draw a potential energy diagram for the reaction showing appropriate labels for both axes,  $E_{\text{a(fwd)}}$ ,  $E_{\text{a(rev)}}$ , and  $\Delta H_{\text{r}}$ .

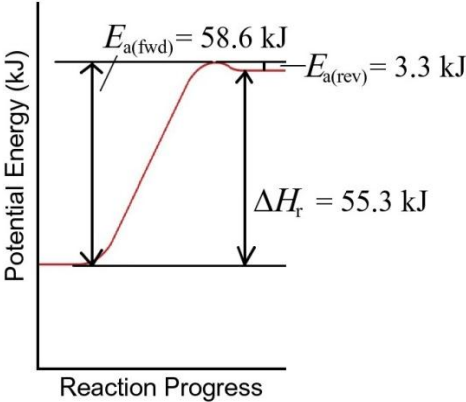
### What Is Required?

You must calculate the activation energy for the reverse reaction and draw a potential energy diagram for the reaction showing appropriate labels for both axes,  $E_{\text{a(fwd)}}$ ,  $E_{\text{a(rev)}}$ , and  $\Delta H_{\text{r}}$ .

### What Is Given?

You know the enthalpy change of the forward reaction:  $\Delta H_{\text{r}} = +55.3 \text{ kJ/mol}$

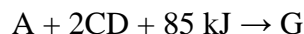
You know the activation energy of the forward reaction:  $E_{\text{a(fwd)}} = +58.6 \text{ kJ}$

Plan Your Strategy	Act on Your Strategy
Use the formula $\Delta H_{\text{r}} = E_{\text{a(fwd)}} - E_{\text{a(rev)}}$ to calculate $E_{\text{a(rev)}}$ .	$\Delta H_{\text{r}} = E_{\text{a(fwd)}} - E_{\text{a(rev)}}$ $+55.3 \text{ kJ} = +58.6 - E_{\text{a(rev)}}$ $E_{\text{a(rev)}} = +3.3 \text{ kJ}$
Draw a labelled energy profile diagram.	 <p>Diagram shows that the reactants have lower potential energy than the products. The activated complex, having the highest potential energy, is positioned between the reactants and products.</p>

### Check Your Solution

Check that the labels are all included. The reaction is endothermic with a small activation energy for the reverse reaction. The potential energy of the products should be higher than the potential energy of the reactants.

20. What is  $E_{a(\text{fwd})}$  for the reaction represented below that has  $E_{a(\text{rev})} = +235 \text{ kJ}$ ?



**What Is Required?**

You need to calculate the activation energy of the forward reaction,  $E_{a(\text{fwd})}$ .

**What Is Given?**

You know the enthalpy change of the forward reaction:  $\Delta H_r = +85 \text{ kJ}$

You know the activation energy of the reverse reaction:  $E_{a(\text{rev})} = +235 \text{ kJ}$

Plan Your Strategy	Act on Your Strategy
Use the formula $\Delta H_r = E_{a(\text{fwd})} - E_{a(\text{rev})}$ to calculate $E_{a(\text{fwd})}$ .	$\Delta H_r = E_{a(\text{fwd})} - E_{a(\text{rev})}$ $+85 \text{ kJ} = E_{a(\text{fwd})} - (+235 \text{ kJ})$ $E_{a(\text{fwd})} = +320 \text{ kJ}$

**Check Your Solution**

Check that the given values have been substituted correctly with the proper sign.