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More Practice Calculating Specific Heat

Lesson 17.1

Step-by-Step Practice

For Problem 1, students are provided a framework and the necessary work for each step. For Problem 2, they are given a framework but must do some of the work themselves.

1. The temperature of a granite rock with a mass of 215 g increases from 24.0°C to 35.5°C when the rock absorbs 1950 J of heat. What is the specific heat of granite?

❶ Analyze **List the knowns and unknown.**

Knowns	Unknown
$m_{granite} = 215\ g$	$C_{granite} = ?\ J/(g \cdot °C)$
$T_f = 35.5°C$	
$T_i = 24.0°C$	
$q = 1950\ J$	

❷ Calculate **Solve for the unknown.**

Before using the equation for specific heat, you must find the change in temperature, ΔT.

$$\Delta T = T_f - T_i = 35.5°C - 24.0°C = 11.5°C$$

Write the equation for specific heat.

$$C_{granite} = \frac{q}{m_{granite} \times \Delta T}$$

Substitute the knowns into the equation and solve.

$$C_{granite} = \frac{1950\ J}{215\ g \times 11.5°C} = 0.789\ J/(g \cdot °C)$$

❸ Evaluate **Does the result make sense?**

The calculated value of 0.789 J/(g·°C) is higher than many metals but lower than the specific heat of water, which is 4.18 J/(g·°C). The result seems reasonable.

2. A 15.0-g block of zinc is immersed in boiling water. The temperature of the block increases from 22.0°C to 100.0°C as it absorbs 459 J of heat. What is the specific heat of zinc?

❶ Analyze **List the knowns and unknown.**

Knowns	Unknown
$m_{zinc} = 15.0\ g$	$C_{zinc} = ?\ J/(g \cdot °C)$
$T_f =$ 100.0°C	
$T_i =$ 22.0°C	
$q =$ 459 J	

2 Calculate **Solve for the unknown.**

Find the change in temperature, ΔT.

$$\Delta T = T_f - T_i$$

$$\Delta T = \underline{100.0°C} - \underline{22.0°C} = \underline{78.0°C}$$

Write the equation for specific heat.

$$C_{zinc} = \frac{q}{m_{zinc} \times \Delta T}$$

Substitute the knowns into the equation and solve.

$$C_{zinc} = \frac{459\ J}{15.0\ g \times 78.0°C} = 0.392\ J/(g \cdot °C)$$

3 Evaluate **Does the result make sense?**

The specific heat of common metals is within the range 0.1 J/(g·°C) to 2.0 J/(g·°C). Is your calculated answer within this range? ___yes___

On Your Own

3. A 6.5-g block of aluminum absorbs 9.0×10^2 J of heat when its temperature increases by 150°C. What is the specific heat of aluminum?

4. A 245-g drinking glass is left outside on a hot day. Its temperature increases from 21.5°C to 31.5°C as it absorbs 2060 J of heat. What is the specific heat of the glass?

5. A musician's brass trumpet has a mass of 1950 g. During an outdoor concert on a hot summer day, the trumpet's temperature increases from 20.5°C to 28.0°C. If the trumpet absorbed 5.37×10^3 J of heat, what is its specific heat?

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More Practice Calculating Enthalpy Change

Lesson 17.2

Step-by-Step Practice

For Problem 1, students are provided a framework and the necessary work for each step. For Problem 2, they are given a framework but must do some of the work themselves.

1. How many kilojoules of heat are produced when 2.30 mol CH_4 burns in excess O_2?

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) + 890\text{ kJ}$$

① Analyze **List the knowns and unknown.**

Knowns	Unknown
amount of CH_4 = 2.30 mol	ΔH = ? kJ for 2.30 mol
ΔH = −890 kJ for 1 mol	

② Calculate **Solve for the unknown.**

Use the thermochemical equation to write a conversion factor that will allow you to convert from moles of CH_4 to kJ of heat.

$$\frac{-890\text{ kJ}}{1\text{ mol } CH_4(g)}$$

Solve for ΔH by multiplying the given amount of CH_4 by the conversion factor.

$$\Delta H = 2.30\ \cancel{\text{mol } CH_4(g)} \times \frac{-890\text{ kJ}}{1\ \cancel{\text{mol } CH_4(g)}}$$
$$= -2.05 \times 10^3\text{ kJ}$$

The reaction produces 2.05×10^3 kJ of heat.

③ Evaluate **Does the result make sense?**

Burning 1 mol $CH_4(g)$ produces 890 kJ of heat. Burning 2 mol $CH_4(g)$ produces 2×890 kJ = 1780 kJ, or 1.78×10^3 kJ. This estimate is close to the calculated answer of 2.05×10^3 kJ.

2. Calculate the amount of heat required to decompose 7.03 mol $CaCO_3(s)$.

$$CaCO_3(s) + 176\text{ kJ} \longrightarrow CaO(s) + CO_2(g)$$

① Analyze **List the knowns and unknown.**

Knowns	Unknown
amount of $CaCO_3$ = 7.03 mol	ΔH = ? kJ for 7.03 mol
ΔH = 176 kJ for 1 mol	

② Calculate **Solve for the unknown.**

Use the thermochemical equation to write a conversion factor that will allow you to convert from moles of $CaCO_3$ to kJ of heat. The units for the conversion factor have been placed for you. Fill in the correct numbers.

$$\frac{176\ \text{kJ}}{1\ \text{mol}\ CaCO_3(s)}$$

Solve for ΔH by multiplying the given amount of $CaCO_3$ by the conversion factor.

$$\Delta H = 7.03\ \cancel{\text{mol}\ CaCO_3(s)} \times \frac{176\ \text{kJ}}{1\ \cancel{\text{mol}\ CaCO_3(s)}}$$
$$= 1.24 \times 10^3\ \text{kJ}$$

③ Evaluate **Does the result make sense?**

The decomposition of 1 mol $CaCO_3(s)$ requires 176 kJ of heat.

Is your calculated answer about 7 times this amount? yes

On Your Own

3. Sulfur reacts with oxygen gas to form sulfur dioxide. The reaction is exothermic. How much energy is released when 4.05 mol S reacts with an excess of O_2?

$S(s) + O_2(g) \longrightarrow SO_2(g) + 297\ \text{kJ}$

4. Calculate the enthalpy change when 2.0 mol NH_3 reacts with an excess of O_2 according to the following equation.

$4NH_3(g) + 7O_2(g) \longrightarrow 4NO_2(g) + 6H_2O(l) + 1131\ \text{kJ}$

5. How many kilojoules of heat are produced by burning 9.1 g of acetylene (C_2H_2)?

$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(l) + 2600\ \text{kJ}$

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More Practice Heat in Changes of State

Lesson 17.3

Step-by-Step Practice

For Problem 1, students are provided a framework and the necessary work for each step. For Problem 2, they are given a framework but must do the work themselves.

1. How many grams of ice will melt at 0°C when 3.23 kJ of heat is added? The molar heat of fusion for water is 6.01 kJ/mol.

1 Analyze **List the knowns and the unknown.**

Knowns	Unknown
$\Delta H_{fus} = 6.01$ kJ/mol	$m_{ice} = ?$ g
$\Delta H = 3.23$ kJ	
$T_i = T_f = 0°C$	

2 Calculate **Solve for the unknown.**

To solve this problem, you need two conversion factors, one that relates heat to moles and one that relates moles to grams.

Express ΔH_{fus} as a conversion factor. Express the molar mass of ice as a conversion factor.

$$\frac{1 \text{ mol } H_2O(s)}{6.01 \text{ kJ}} \qquad \frac{18.0 \text{ g } H_2O(s)}{1 \text{ mol } H_2O(s)}$$

Multiply the amount of heat added by both conversion factors. Then solve the equation.

$$m_{ice} = \frac{3.23 \cancel{\text{kJ}}\ 1 \cancel{\text{mol } H_2O(s)}}{6.01 \cancel{\text{kJ}}} \times \frac{18.0 \text{ g } H_2O(s)}{1 \cancel{\text{mol } H_2O(s)}}$$

$$m_{ice} = 9.67 \text{ g}$$

3 Evaluate **Does the result make sense?**

To melt 1 mol of ice, 6.01 kJ of energy is needed. About half this amount of heat is available, so about 0.5 mol of ice, or about 9 g, should melt.

2. How many kilojoules of heat are absorbed when 23.8 g of ethanol vaporizes at its normal boiling point, 78.3°C? For ethanol, the molar heat of vaporization is 0.90 kJ/mol. The molar mass of ethanol is 46.07 g/mol.

1 Analyze **List the knowns and the unknown.**

Knowns	Unknown
ΔH_{vap} = 0.90 kJ/mol	$\Delta H = ?$ kJ
$m_{ethanol}$ = 23.8 g	
molar mass = 46.07 g/mol	

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② Calculate **Solve for the unknown.**

Use the molar mass of ethanol to write a conversion factor that will convert grams of ethanol to moles of ethanol. You also need to express ΔH_{vap} as a conversion factor.

$$\frac{1 \text{ mol } C_2H_6O(l)}{46.07 \text{ g } C_2H_6O(l)} \qquad \frac{0.90 \text{ kJ}}{1 \text{ mol } C_2H_6O(l)}$$

Multiply the mass of ethanol in grams by both conversion factors.

$$\Delta H = 23.8 \cancel{\text{ g } C_2H_6O(l)} \times \frac{1 \cancel{\text{ mol } C_2H_6O(l)}}{46.07 \cancel{\text{ g } C_2H_6O(l)}} \times \frac{0.90 \text{ kJ}}{1 \cancel{\text{ mol } C_2H_6O(l)}}$$

$\Delta H =$ 0.46 kJ

③ Evaluate **Does the result make sense?**

The mass of ethanol (23.75 g) is about half the molar mass of ethanol (46.07 g). So it makes sense that the amount of heat added (0.46 kJ) is about half of 0.90 kJ.

On Your Own

3. How many grams of ice at 0°C will melt if you add 2.18 kJ of heat? The molar heat of fusion for water is 6.01 kJ/mol.

4. Calculate the enthalpy change when 1.26 g of oxygen gas condenses at its normal boiling point. The molar heat of vaporization for oxygen is 6.82 kJ/mol.

5. A beaker contains 5.80 g of ice. If you add 0.615 kJ of heat, how many grams of the ice will stay frozen? *Hint:* First find how many grams will melt.