

## Practice Test – Chapter 5 - Thermochemistry

Goal: I can give examples of different forms of energy. I can also state common units of energy and convert between these units.

1. Choose the correct type of energy for each of the following: Choices: potential energy, kinetic energy, thermal energy, internal energy, electrostatic potential energy

Kinetic Energy a) The energy of motion.

Potential Energy b) The energy of an object due to its relative position.

Electrostatic PE c) The energy due to the attractions or repulsions of charged particles such as protons and electrons.

Thermal Energy d) The energy that a substance possess because of its temperature.

Internal Energy e) The sum of all the PE and KE of a system.

2. Circle all of the following which are units of energy:

joule kilogram gram calorie Calorie kilojoule kg•m<sup>2</sup>/s<sup>2</sup> m<sup>2</sup>/s<sup>2</sup>

same as a Joule

Goal: I can define the first law of thermodynamics and write the associated equation.

3. The First Law of Thermodynamics states that ... Energy is conserved

Energy lost by system is gained by surroundings

(and vice-versa)

4. Which equation represents the First Law of Thermodynamics?

- a.  $\Delta E = q + w$   
 b.  $\Delta H = q_p$   
 c.  $q = m \cdot C_p \cdot \Delta T$   
 d.  $KE = \frac{1}{2} mv^2$

Goal: I can describe how the change in internal energy of a system is related to the changes of heat and work between the system and its surroundings.

5. A positive "q" means that the system \_\_\_\_\_ heat. A positive "w" means that work is done \_\_\_\_\_ the system.

a. loses, by

b. loses, on

c. gains, by

d. gains, on



6. Given:  
 $\Delta E = ?$   
 $w = -230 \text{ kJ}$   
 $q = -130 \text{ kJ}$

Soln:  
 $\Delta E = q + w$   
 $\Delta E = -130 \text{ kJ} + -230 \text{ kJ}$   
 $\Delta E = -360 \text{ kJ}$

6. What is  $\Delta E$  for a system when it does 230 kJ of work on its surroundings and 130 kJ of heat is removed from the system?  
 a. +100 kJ      b. -100 kJ      c. +360 kJ      **d. -360 kJ**

Goal: I can define enthalpy and relate the enthalpy change in a process to the heat added (endothermic) to or lost (exothermic) by the system during the process.

7. Enthalpy is the same thing as heat when measured at a constant \_\_\_\_\_.  
 a. volume      b. temperature      **c. pressure**      d. elevation

8. A negative  $\Delta H$  indicates that a reaction is ... **Exothermic**  
 a. endothermic.  
 b. absorbing heat.  
 c. occurring spontaneously.  
 d. nonspontaneous.  
**e. losing heat.**

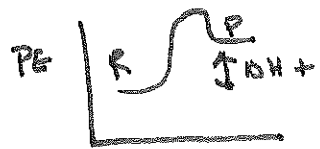


9. Assume that a chemical reaction is occurring in a plastic bag and that you are holding the bag in your hands. If your hands feel cold it is because ...  
 a. the chemical reaction is exothermic and absorbing heat from your hands.  
**b. the chemical reaction is endothermic and absorbing heat from your hands.**  
 c. the chemical reaction is exothermic and releasing heat into your hands.  
 d. the chemical reaction is endothermic and releasing heat into your hands.

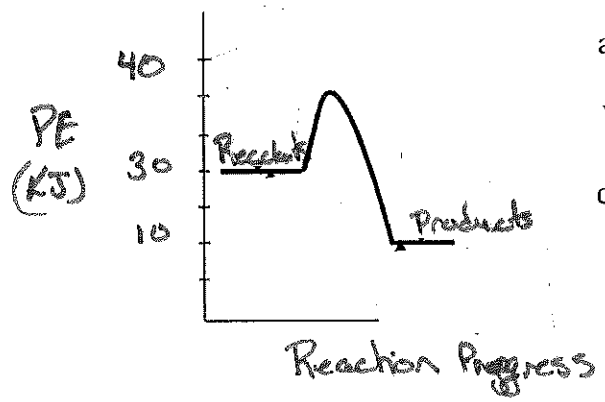
Goal: I can differentiate between exothermic and endothermic processes. I can also sketch an exothermic and endothermic energy diagram and label the various parts.

10. Which ONE statement concerning the following is correct?  
 $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{NO}(\text{g}); \Delta H = +43.2 \text{ kcal}$

- a. The reaction is exothermic.  
 b. The products have an enthalpy loss.  
 c. The sign of  $\Delta H$  for the reaction is negative.  
**d. PE of the products exceeds that of the reactants.**  
 e. The products have less enthalpy than the reactants.



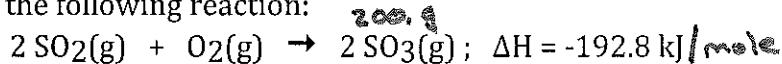
11. Consider the following energy diagram:  
 $\text{A} + \text{B} \rightarrow \text{C} + \text{D}$



- a. Is the reaction **exothermic** or endothermic?  
 b. Label the reactants and products  
 c. What is the value of  $\Delta H$ ? **-20 kJ**

Goal: I can make stoichiometric calculations based upon a thermochemical equation.

12. Consider the following reaction:



Assume enough  $\text{SO}_2$  and  $\text{O}_2$  were reacted and 200. grams of  $\text{SO}_3$  were produced. Calculate the corresponding amount of heat produced.

- a. -3.86 kJ    b. -15.1 kJ    c. -75.5 kJ    d. -241 kJ    e. -521 kJ

5  $32.06 = 32.06$   
 16.01 x 3  $\frac{48.03}{80.09}$

$$\left( \frac{200. \text{g SO}_3}{1} \right) \left( \frac{1 \text{ mole SO}_3}{80.09 \text{g SO}_3} \right) \left( \frac{-192.8 \text{ kJ}}{2 \text{ mole SO}_3} \right) = -240 \text{ kJ}$$

13. Typical heat capacity units are \_\_\_\_\_ and specific heat capacity units are \_\_\_\_\_.
- a.  $\text{J}^\circ\text{C}$ ,  $\text{J/g}^\circ\text{C}$     b.  $\text{J/g}^\circ\text{C}$ ,  $\text{J}^\circ\text{C}$     c.  $^\circ\text{C/J}$ ,  $^\circ\text{C} \cdot \text{J/g}$     d.  $^\circ\text{C} \cdot \text{J/g}$ ,  $^\circ\text{C/J}$

Goal: I can solve problems using  $q = m \cdot C_p \cdot \Delta T$ ; calculate any one of the quantities given the other three.

14. What is the final temperature of 30. g of Al if 540. J of heat is added to a sample at  $25.0^\circ\text{C}$ ? (specific heat of Al =  $0.90 \text{ J/g}^\circ\text{C}$ )

- a.  $10^\circ\text{C}$     b.  $20^\circ\text{C}$     c.  $35^\circ\text{C}$     d.  $45^\circ\text{C}$     e.  $135^\circ\text{C}$

Soln:  $q = m \cdot C_p \cdot \Delta T$   
 $\Delta T = \frac{q}{m \cdot C_p} = \frac{540. \text{J}}{(30. \text{g})(0.90 \text{ J/g}^\circ\text{C})} = 20^\circ\text{C}$   
 $\Delta T = T_f - T_i$   
 $T_f = \Delta T + T_i = 20^\circ\text{C} + 25.0^\circ\text{C} = 45.0^\circ\text{C}$

Given:  
 $m = 30. \text{g}$   
 $C_p = 0.90 \text{ J/g}^\circ\text{C}$   
 $T_i = 25.0^\circ\text{C}$   
 $q = 540. \text{J}$   
 $T_f = ?$

15. Assume that 100.0 g of water  $30.0^\circ\text{C}$  is mixed with 50.0 g of water at  $0.0^\circ\text{C}$ . What is the final temperature?

- a.  $40.0^\circ\text{C}$     b.  $20.0^\circ\text{C}$     c.  $15.0^\circ\text{C}$     d.  $10.0^\circ\text{C}$

16. How much heat is lost when 10.0 g of iron cools from  $129^\circ\text{C}$  to  $79^\circ\text{C}$ ? (The specific heat of iron is  $0.450 \text{ J/g}^\circ\text{C}$ .)

- a. 12.0 J    b. 24 J    c. 124 J    d. 175 J    e. 225 J

Given:  
 $m = 10.0 \text{g}$   
 $q = -? \text{J}$   
 $C_p = 0.450 \text{ J/g}^\circ\text{C}$   
 $\Delta T = 79^\circ - 129^\circ\text{C}$   
 $\Delta T = -50^\circ\text{C}$   
 Soln:  
 $q = m \cdot C_p \cdot \Delta T$   
 $= (10.0 \text{g})(0.450 \frac{\text{J}}{\text{g}^\circ\text{C}})(-50^\circ\text{C})$   
 $q = -225 \text{ J}$   
 or  
 $q = 225 \text{ J}$   
Lost

15.

Given:  
 $m_H = 100.0 \text{g}$   
 $T_H = 30.0^\circ\text{C}$   
 $m_c = 50.0 \text{g}$   
 $T_c = 0.0^\circ\text{C}$

Soln:  
 $-q \text{ lost by Hot H}_2\text{O} = q \text{ gained by cold H}_2\text{O}$   
 $-(m_H) C_p \Delta T_H = m_c C_p \Delta T_c$   
 $-100.0 (T_f - 30^\circ\text{C}) = 50.0 \text{g} (T_f - 0^\circ\text{C})$   
 $-100 T_f + 3000.^\circ\text{C} = 50.0 T_f - 0$   
 $3000^\circ\text{C} = 150 T_f$   
 $T_f = 20^\circ\text{C}$

17. Consider the following specific heats of metals:

<u>Metal</u>	<u>Specific Heat (J/g°C)</u>
Lithium	3.56
Gallium	0.372
Nickel	0.444
Gold	0.129
Sodium	1.23

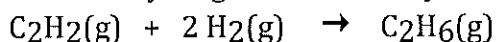
If the same amount of heat is added to 200 g samples of each of the metals, assume all metals are at the same temperature, which metal will attain the lowest temperature?

- ← highest Cp!
- a. Lithium    b. Gallium    c. Nickel    d. Gold    e. Sodium

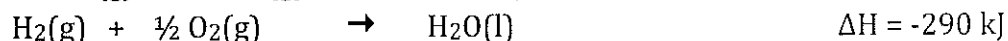
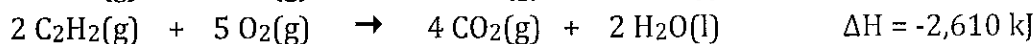
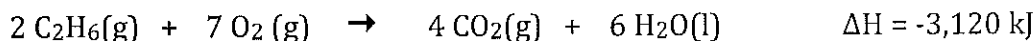
Goal: I can state Hess's law and use Hess's Law to solve problems.

18. State Hess's Law. *The heat change ( $\Delta H$ ) in a chemical rxn is the same regardless of the number of stages/series of rxn's*

19. What is the heat of hydrogenation of acetylene, at 25°C and 1 atm,



given the following thermochemical equations:



- a. -325 kJ    b. +325 kJ    c. +1,610 kJ    d. -1,610 kJ

*work:  
on separate  
piece of  
paper*

Goal: I can define the terms *standard state and standard heat of formation*. I will also be to write a chemical reaction associated with the standard heat of formation.

20. Which of the following conditions represent standard states?

- a. 1 atm & 0°C    b. 2 atm & 0°C    c. 2 atm & 273°C    d. 1 atm & 25°C

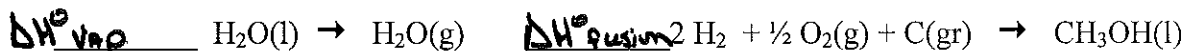
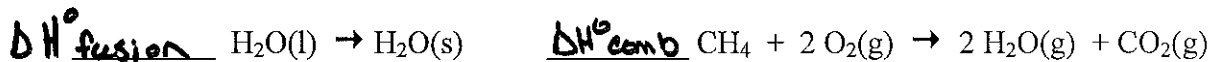
21. Which equation represents  $\Delta H_f^\circ$  for  $\text{NO}_2$ ?

- a.  $\text{NO}(\text{g}) + \text{O}(\text{g}) \rightarrow \text{NO}_2(\text{g})$   
 b.  $\text{NO}(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g})$   
 c.  $2 \text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{NO}_2(\text{g})$   
 d.  $\text{N}_2(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow 2 \text{NO}_2(\text{g})$   
 e.  $\frac{1}{2} \text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g})$

*$\Delta H_f^\circ$  = formation!  
must make 1 mole  $\text{NO}_2$*

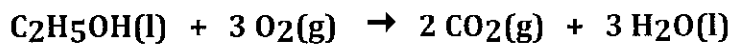
22. Choose the proper  $\Delta H$  for each equation below. Place the correct  $\Delta H$  on the provided blanks.

**CHOICES:**  $\Delta H^\circ_{\text{fusion}}$   $\Delta H^\circ_{\text{vap}}$   $\Delta H^\circ_f$   $\Delta H^\circ_{\text{comb}}$



Goal: I can calculate the enthalpy change in a reaction when given the standard enthalpies of formations of each reactant and product.

23. What is  $\Delta H^\circ_f$  for one mole of  $\text{C}_2\text{H}_5\text{OH}(\text{l})$  given the following data at  $25^\circ\text{C}$ :



$$\Delta H^\circ_{\text{reaction}} = -1,366 \text{ kJ/mole rxn}$$

$$\Delta H^\circ_f \text{ for } \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$$

$$\Delta H^\circ_f \text{ for } \text{H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol}$$

- a. +462.6 kJ      b. +278.4 kJ      c. -462.6 kJ      d. 278.4 kJ

24. What is the heat of combustion for one mole of benzene at  $25^\circ\text{C}$  and 1 atm,



given the following data:

$$\Delta H^\circ_f \text{ for } \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$$

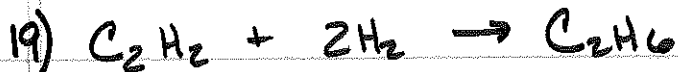
$$\Delta H^\circ_f \text{ for } \text{H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol}$$

$$\Delta H^\circ_f \text{ for } \text{C}_6\text{H}_6(\text{l}) = 49.0 \text{ kJ/mol}$$

- a. +3,271 kJ      b. -3,271 kJ      c. -636 kJ      d. +636 kJ

Unity - Practice Test - MC

3/5



$\Delta H_{rxn} = -1366 \text{ KJ}$

$\Delta H_{rxn} = \sum \Delta H^\circ_{products} - \sum \Delta H^\circ_{reactants}$

$\sum \Delta H^\circ_{reactants} = \sum \Delta H^\circ_{products} - \Delta H_{rxn}$

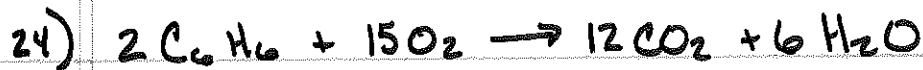
$\Delta H^\circ_{C_2H_5OH} + 0 = \left[ \frac{2 \text{ mole}}{\text{mole rxn}} (-393 \frac{\text{KJ}}{\text{mole}}) + \frac{3 \text{ mole}}{\text{mole rxn}} (-285.8 \frac{\text{KJ}}{\text{mole}}) \right] - (-1366 \frac{\text{KJ}}{\text{mole rxn}})$

$= -786 \frac{\text{KJ}}{\text{mole rxn}} + (-857.4 \frac{\text{KJ}}{\text{mole rxn}}) + 1366 \frac{\text{KJ}}{\text{mole rxn}}$

$\Delta H^\circ_{C_2H_5OH} = -277.4 \frac{\text{KJ}}{\text{mole rxn}}$

← Net this is for 2 moles

2/5



$$\Delta H_f^\circ = \quad 49.0 \text{ kJ/mol} \quad 0 \quad -393.5 \text{ kJ/mol} \quad -285.8 \text{ kJ/mol}$$

$$\Delta H_{\text{comb}}^\circ \text{C}_6\text{H}_6 = ?$$

$$\Delta H_{\text{comb}}^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$= \left( \frac{12 \text{ mol}}{1 \text{ mol rxn}} \right) (-393.5 \frac{\text{kJ}}{\text{mol}}) + \frac{6 \text{ mol}}{1 \text{ mol rxn}} (-285.8 \frac{\text{kJ}}{\text{mol}}) - \left[ \frac{2 \text{ mol}}{1 \text{ mol rxn}} (49.0 \frac{\text{kJ}}{\text{mol}}) \right]$$

$$= -4722 \frac{\text{kJ}}{\text{mol rxn}} - 1715 \frac{\text{kJ}}{\text{mol rxn}} - (98.0 \frac{\text{kJ}}{\text{mol rxn}})$$

$$= -6535 \text{ kJ/mol rxn}$$

2 mol

$$\Delta H_{\text{comb}}^\circ = -3268 \text{ kJ/mol rxn}$$

① A) Equipment needed: goggles, Thermometer, Calorimeter, graduated cylinder

② measurements: initial temp of Acids & Bases  
final Temp of Soln (Acid & Base)  
Volumes of Acid & Base

③  $q = m c_p \Delta T$  multiple mass <sup>of soln</sup> By specific heat of Soln By the change in Temp of Soln.

Divide the Product By moles of either Acid or Base

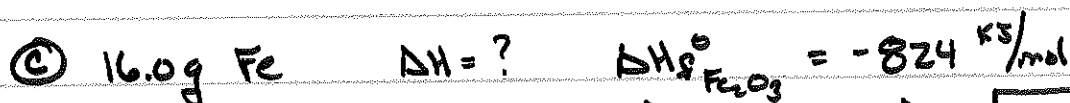
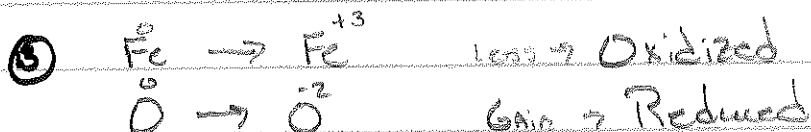
④ Possible errors for lower value

① loss of heat to the calorimeter, air, or to Thermometer

② incomplete transfer of either the Acid or Base (not all poured out of grad cylinder) or spilling of Acid and or Base



②



$$\left( \frac{16.0 \text{g Fe}}{55.85 \text{g Fe}} \right) \left( \frac{2 \text{ mole Fe}_2\text{O}_3}{4 \text{ mole Fe}} \right) \left( \frac{-824 \text{ kJ}}{1 \text{ mole Fe}_2\text{O}_3} \right) = \boxed{-118 \text{ kJ}}$$

③ Given

Al

$$T_1 = 180.0^\circ\text{C}$$

m = ? g Al

$$C_p = 0.941 \text{ J/g}^\circ\text{C}$$

H<sub>2</sub>O

$$T_1 = 19.0^\circ$$

$$T_2 = 25.0^\circ\text{C}$$

$$m = 400.0 \text{g}$$

$$C_p = 4.18 \text{ J/g}^\circ\text{C}$$

$$\Delta T_w = T_2 - T_1 = 25.0^\circ - 19.0^\circ = 7.0^\circ$$

Sol:Final Temp of Al equal to final Temp H<sub>2</sub>O = 25.0°C

$$\text{Al: } \Delta T_{\text{Al}} = 25.0^\circ\text{C} - 180.0^\circ\text{C} = -155.0^\circ\text{C}$$

$$q_{\text{gain}} \text{ By H}_2\text{O} = -q_{\text{loss}} \text{ By Al}$$

$$m_w C_{pw} \Delta T_w = -m_{\text{Al}} C_{p\text{Al}} \Delta T_{\text{Al}}$$

$$(400.0 \text{g})(4.18 \text{ J/g}^\circ\text{C})(7.0^\circ\text{C}) = -m_{\text{Al}}$$

$$(0.941 \text{ J/g}^\circ\text{C})(-155.0^\circ\text{C})$$

$$-m_{\text{Al}} = -80.0^\circ\text{C}$$

$$\boxed{m_{\text{Al}} = 80.0^\circ\text{C}}$$

AP Chem - Unit 4 - Practice Test - FR

5/9

Given:

④  $\text{N}_2\text{H}_4$   
 $m = 1.0 \text{ g}$   
 $C_p = 840. \text{ J/}^\circ\text{C}$

$\text{H}_2\text{O}$   
 $m = 1200. \text{ g}$   
 $T_1 = 24.62^\circ\text{C}$   
 $T_2 = 28.16^\circ\text{C}$   
 $C_p = 4.18 \text{ J/g}^\circ\text{C}$   
 $\Delta T_{\text{H}_2\text{O}} = 28.16 - 24.62 = 3.54^\circ\text{C}$

$\Delta H^\circ_{\text{comb. N}_2\text{H}_4} = ?$

Soln:

$$q_{\text{rxn}} = -(\text{heat absorbed by H}_2\text{O} + \text{heat absorbed by calorimeter})$$

$$= m_{\text{H}_2\text{O}} C_{p,\text{H}_2\text{O}} \Delta T_{\text{H}_2\text{O}} + (C_{\text{calorimeter}}) \Delta T_{\text{cal}}$$

$$= -[(1200. \text{ g})(4.18 \frac{\text{J}}{\text{g}^\circ\text{C}})(3.54^\circ\text{C}) + (840. \text{ J/}^\circ\text{C})(3.54^\circ\text{C})]$$

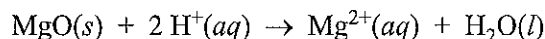
$$= -(20730.24 \text{ J})$$

$q_{\text{rxn}} = -20700 \text{ J}$

= sign  
 heat Released By  
 Combustion Process

**AP<sup>®</sup> CHEMISTRY**  
**2013 SCORING GUIDELINES**

**Question 3**  
**(9 points)**



A student was assigned the task of determining the enthalpy change for the reaction between solid MgO and aqueous HCl represented by the net-ionic equation above. The student uses a polystyrene cup calorimeter and performs four trials. Data for each trial are shown in the table below.

Trial	Volume of 1.0 M HCl (mL)	Mass of MgO(s) Added (g)	Initial Temperature of Solution (°C)	Final Temperature of Solution (°C)
1	100.0	0.25	25.5	26.5
2	100.0	0.50	25.0	29.1
3	100.0	0.25	26.0	28.1
4	100.0	0.50	24.1	28.1

$\Delta T$        $\Delta T / \text{mass}$

1.0 /      4

4.1 ✓      8.2

2.6 ✓      9.4

4.1      8.2

(a) Which is the limiting reactant in all four trials, HCl or MgO? Justify your answer.

$0.100 \text{ L} \times \frac{1.0 \text{ mol HCl}}{1.0 \text{ L}} = 0.10 \text{ mol HCl}$ $0.50 \text{ g MgO} \times \frac{1 \text{ mol MgO}}{40.30 \text{ g MgO}} = 0.0124 \text{ mol MgO}$ <p>By the stoichiometry of the equation, only <math>2 \times (0.0124 \text{ mol}) = 0.0248 \text{ mol HCl}</math> is needed to react with the MgO, thus HCl is in excess and MgO is limiting.</p> <p>OR</p> <p>The temperature change depended on the amount of MgO added, indicating that MgO was the limiting reactant.</p>	<p>1 point is earned for the correct choice with justification.</p>
--	---

(b) The data in one of the trials is inconsistent with the data in the other three trials. Identify the trial with inconsistent data and draw a line through the data from that trial in the table above. Explain how you identified the inconsistent data.

<p>Trial 1 is inconsistent.</p> <p>The temperature change should be directly proportional (approximately) to the amount of the limiting reactant present. The ratio <math>\Delta T / (\text{mass MgO})</math> should be constant. In trial 1, the ratio is one-half of trials 2, 3, and 4. Therefore, trial 1 is inconsistent with the other trials.</p>	<p>1 point is earned for identifying trial 1 with a valid justification.</p>
--	--

**AP<sup>®</sup> CHEMISTRY**  
**2013 SCORING GUIDELINES**

**Question 3 (continued)**

For parts (c) and (d), use the data from one of the other three trials (i.e., not from the trial you identified in part (b) above). Assume the calorimeter has a negligible heat capacity and that the specific heat of the contents of the calorimeter is  $4.18 \text{ J/(g}\cdot\text{C}^\circ)$ . Assume that the density of the  $\text{HCl(aq)}$  is  $1.0 \text{ g/mL}$ .

- (c) Calculate the magnitude of  $q$ , the thermal energy change, when the  $\text{MgO}$  was added to the  $1.0 \text{ M}$   $\text{HCl(aq)}$ . Include units with your answer.

<p style="text-align: center;"><math>q_{\text{calorimeter}} = q_{\text{cal}} = mc\Delta T</math></p> <p>In trial 2, <math>q_{\text{cal}} = \left[ \left( 100.0 \text{ mL} \times \frac{1.0 \text{ g}}{\text{mL}} \right) + 0.50 \text{ g} \right] \left( \frac{4.18 \text{ J}}{\text{g}\cdot\text{C}^\circ} \right) (4.1^\circ\text{C}) = 1700 \text{ J or } 1.7 \text{ kJ}</math></p> <p>OR</p> <p>In trial 3, <math>q_{\text{cal}} = \left[ \left( 100.0 \text{ mL} \times \frac{1.0 \text{ g}}{\text{mL}} \right) + 0.25 \text{ g} \right] \left( \frac{4.18 \text{ J}}{\text{g}\cdot\text{C}^\circ} \right) (2.1^\circ\text{C}) = 880 \text{ J or } 0.88 \text{ kJ}</math></p> <p>OR</p> <p>In trial 4, <math>q_{\text{cal}} = \left[ \left( 100.0 \text{ mL} \times \frac{1.0 \text{ g}}{\text{mL}} \right) + 0.50 \text{ g} \right] \left( \frac{4.18 \text{ J}}{\text{g}\cdot\text{C}^\circ} \right) (4.0^\circ\text{C}) = 1700 \text{ J or } 1.7 \text{ kJ}</math></p>	<p>1 point is earned for the correct mass of the solution.</p> <p>1 point is earned for the correct calculation of <math>q</math> for any trial with a valid <math>\Delta T</math> and correct units.</p>
--	---

- (d) Determine the student's experimental value of  $\Delta H^\circ$  for the reaction between  $\text{MgO}$  and  $\text{HCl}$  in units of  $\text{kJ/mol}_{\text{rxn}}$ .

<p>Assuming that no heat was lost to the surroundings, <math>q_{\text{rxn}} = -q_{\text{cal}}</math>.</p> <p>In trials 2 and 4,</p> $\Delta H^\circ = \frac{q_{\text{rxn}}}{n_{\text{MgO}}} = \frac{-1,700 \text{ J}}{0.50 \text{ g MgO} \times \frac{1 \text{ mol MgO}}{40.30 \text{ g MgO}}} = -140,000 \text{ J/mol}_{\text{rxn}} \times \frac{1 \text{ kJ}}{1000 \text{ J}}$ <p style="margin-left: 20px;"><math>= -140 \text{ kJ/mol}_{\text{rxn}}</math></p> <p>In trial 3,</p> $\Delta H^\circ = \frac{-880 \text{ J}}{0.25 \text{ g MgO} \times \frac{1 \text{ mol MgO}}{40.30 \text{ g MgO}}} = -140,000 \text{ J/mol}_{\text{rxn}} \times \frac{1 \text{ kJ}}{1000 \text{ J}}$ <p style="margin-left: 20px;"><math>= -140 \text{ kJ/mol}_{\text{rxn}}</math></p>	<p>1 point is earned for the correct calculation of moles of <math>\text{MgO}</math> or setup of equation.</p> <p>1 point is earned for the value of <math>\Delta H^\circ</math> and sign consistent with the setup.</p>
--	--

$\left( \frac{0.50 \text{ g MgO}}{40.30} \right) \times 1 \text{ mole} = 0.124 \text{ moles}$   
 $\Delta H = -q$   
 $\frac{1700 \text{ kJ}}{1}$

$q = M_{\text{rxn}} C_{\text{p,rxn}} \Delta T = M_{\text{sol}} C_{\text{p,sol}} \Delta T$

**AP<sup>®</sup> CHEMISTRY**  
**2013 SCORING GUIDELINES**

**Question 3 (continued)**

- (e) Enthalpies of formation for substances involved in the reaction are shown in the table below. Using the information in the table, determine the accepted value of  $\Delta H^\circ$  for the reaction between  $\text{MgO}(s)$  and  $\text{HCl}(aq)$ .

Substance	$\Delta H_f^\circ$ (kJ/mol)
$\text{MgO}(s)$	-602
$\text{H}_2\text{O}(l)$	-286
$\text{H}^+(aq)$	0
$\text{Mg}^{2+}(aq)$	-467

$\Delta H^\circ = \sum n_p \Delta H_f^\circ \text{ products} - \sum n_r \Delta H_f^\circ \text{ reactants}$ $= [\Delta H_f^\circ \text{Mg}^{2+}(aq) + \Delta H_f^\circ \text{H}_2\text{O}(l)] - [\Delta H_f^\circ \text{MgO}(s) + 2 \Delta H_f^\circ \text{H}^+(aq)]$ $= [-467 \text{ kJ/mol} + (-286 \text{ kJ/mol})] - [-602 \text{ kJ/mol} + 2(0) \text{ kJ/mol}]$ $= -151 \text{ kJ/mol}_{\text{rxn}}$	<p>1 point is earned for the correct setup using the <math>\Delta H_f^\circ</math> values.</p> <p>1 point is earned for the correct value and sign consistent with the setup.</p>
--	---

- (f) The accepted value and the experimental value do not agree. If the calorimeter leaked heat energy to the environment, would it help account for the discrepancy between the values? Explain.

<p>Yes. The experimentally determined value for <math>\Delta H^\circ</math> was less negative than the accepted value. If heat had leaked out of the calorimeter, then the <math>\Delta T</math> of the contents would be less than expected, leading to a smaller calculated value for <math>q</math> and a less negative value for <math>\Delta H^\circ</math>.</p>	<p>1 point is earned for the correct response with a valid explanation.</p>
---	---