H Chapter 16 Practice Problems 2015-2016

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1. Energy is the ability to do work or produce heat
2. Heat flows from warm to cold temps.
3. Heat is the total energy of molecular motion, dependent upon amount, size, and type of particles. Heat is energy.
4. Exothermic process is one that releases heat to its surroundings (feels warm). Endothermic process is one that absorbs heat from the surroundings (feels cold).
5. Specific heat of a substance is the amount of heat required to raise the temperature of one gram of that substance by one degree Celsius.
6. \( Q = m(\Delta T)C_p \)
7. Temperature is a measurement of kinetic energy of heat. Heat is energy.
8. 142,000 cal or 1.42 \( \times \) 10^5; 594,000 J
9. 20.7 kcal or 20.7 Calories
10. 0.0612 kcal
11. 527,000 J
12. 109 kcal
13. Q is positive the reaction is endothermic.
14. a. endothermic because, b. positive, heat is on the reactant side
15. a. exothermic, b. negative, heat is on the product side
16. heat released
17. -5985 kJ
18. \( m = 153.9 \) g ethanol
19. 12.5 \( \bar{J} \)
20. \( \Delta H_{rxn} = \) negative, means it is exothermic
21. \( \Delta H_{rxn} = 66.36 \) kJ, endothermic
22. \( \Delta H_{rxn} = -1397.8 \) kJ, exothermic
23. \( \Delta H_{rxn} = -241.8 \) kJ, exothermic
24. \( \Delta H_{rxn} = -827.23 \), exothermic
25. 21087.36 J; 102 g
26. -38750.4 J; 168 g
27. 58.5°C
28. 6276 J; 0.136 J/g°C; tungsten
29. During a phase change, the temperature stays the same.
30. \( \Delta H_{vap} = -\Delta H_{cond}; \Delta H_{fus} = -\Delta H_{solid} \)
31. You use \( Q = m(\Delta T)C_p \) to calculate energy when there is a temperature change. You use latent heat to calculate energy when there is a phase change.
32. see drawing below
33. 0.493 kJ
34. \( NH_3 \)
35. 144 kJ
8. $142 \text{ Calorie} / 1000 \text{ cal} = 142,000 \text{ cal} / 1 \text{ Calorie}$

$142,000 \text{ cal} / 4.184 \text{ J} = 59,400 \text{ J}$

9. $86.5 \text{ KJ} / 1000 \text{ J} / 1 \text{ cal} / 1 \text{ Kcal} = 20.7 \text{ kcal} = 20.7 \text{ Calories}$

10. $2560 \text{ J} / 1 \text{ cal} / 1 \text{ Kcal} = 0.0612 \text{ kcal} / 4.184 \text{ J} / 1000 \text{ cal}$

11. $126 \text{ Calories} / 1000 \text{ cal} / 4.184 \text{ J} = 5.27,000 \text{ J} (3st)$

12. $455 \text{ KJ} / 1000 \text{ J} / 1 \text{ Cal} / 1 \text{ Kcal} = 109 \text{ kcal} / 1 \text{ KJ} / 4.184 \text{ J} / 1000 \text{ cal}$

17. $250.0 \text{ g C}_8\text{H}_8 / 1 \text{ mol C}_8\text{H}_8 / -5471 \text{ KJ} / -5985 \text{ KJ}$ or

$114.26 \text{ g C}_8\text{H}_8 / 2 \text{ mol C}_8\text{H}_8 / -5985 \text{ KJ}$ released

$7.22 \text{ mL} / -5.58 \text{ J} = -5.58 \text{ J}$
18. \[ \text{C}_2\text{H}_5\text{OH} + 2\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O} + 1367\text{KJ} \]

\[ -4565\text{KJ} \]

\[ 1\text{ mol C}_2\text{H}_5\text{OH} \]

\[ 46.08\text{ g C}_2\text{H}_5\text{OH} = 153.9\text{ g C}_2\text{H}_5\text{OH} \]

\[ -1367\text{KJ} \]

19. \[ \text{C} + 2\text{S} \rightarrow \text{CS}_2 \quad \Delta H = 89.3\text{KJ} \]

\[ 10.67\text{ g CS}_2 \]

\[ 1\text{ mol CS}_2 \]

\[ 89.3\text{KJ} = 12.5\text{KJ} \]

\[ 76.15\text{ g CS}_2 \]

\[ 1\text{ mol CS}_2 \]

\[ 10000.00\text{KJ} \]

21. \[ \text{N}_2 + 2\text{O}_2 \rightarrow 2\text{NO}_2 \quad \Sigma\text{Prod} - \Sigma\text{React.} \]

\[ 33.18 \quad 2(33.18) \]

\[ 0 \quad 2(0+6) \]

\[ 64.36\text{KJ, endo} \]

22. \[ 4\text{NH}_3 + 7\text{O}_2 \rightarrow 4\text{NO}_2 + 6\text{H}_2\text{O} \]

\[ 4(-40.11) + 7(0) \quad 4(33.18) + 6(-285.830) \]

\[ \Sigma\text{Prod} - \Sigma\text{React.} \]

\[ [4(33.18) + 6(-285.830)] - [4(-40.11)] \]

\[ [132.72 + -1714.98] - [-184.44] \]

\[ -1582.26 + 184.44 \]

\[ -1397.82\text{KJ, EXO} \]

23. \[ \text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2 \]

\[ -824.2 + 3(-110.525) \]

\[ 2\cdot 0 \quad 3(-393.509) \]

\[ \Sigma\text{Prod} - \Sigma\text{React.} \]

\[ [3(-393.509)] - [-824.2 + 3(-110.525)] \]

\[ -1180.527 - [-824.2 + -331.575] \]

\[ -1180.527 - [-1155.775] \]

\[ -24.752 \times -24.8\text{KJ} = \text{EXO} \]
24. \[ 2 \text{H}_2\text{S} + 3\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{SO}_2 \]  
\[ \begin{array}{c}
\text{E}_\text{Prod} - \text{E}_\text{React} \\
\left[ 2(-285.830) + 296.830 \right] - \left[ 2(-20.63) + 3 \cdot 0 \right] \\
-868.49 - [-42.26] \\
-868.49 + 41.26 \\
-827.23 \text{kJ, exo}
\end{array} \]

25. \[ q = \int \]  
\[ -mC\Delta T = mC\Delta T \]  
\[ -m(0.449 \text{J/g}^\circ \text{C})(42.0^\circ \text{C} - 50^\circ \text{C}) = 240.9 \text{g}(4.184 \text{J/g}^\circ \text{C})(42.0^\circ \text{C} - 21.0^\circ \text{C}) \]  
\[ m(206.091000) = 21087.3600 \]  
\[ 206.091000 \quad 206.091000 \]  
\[ m \approx 102.9 \text{g Fe} \]

26. \[ -mC\Delta T = mC\Delta T \]  
\[ -135 \text{g}(0.897 \text{J/g}^\circ \text{C})(80.0^\circ \text{C} - 400.0^\circ \text{C}) = m(4.184 \text{J/g}^\circ \text{C})(80.0^\circ \text{C} - 25.0^\circ \text{C}) \]  
\[ 38750.4 = m(230.12) \]  
\[ \frac{38750.4}{230.12} = 230.12 \]  
\[ 16.8 \text{g H}_2\text{O} = m \]

27. \[ q = \int \]  
\[ -9750 \text{J} = 335 \text{g}(4.184 \text{J/g}^\circ \text{C})(T_F - 65.5^\circ \text{C}) \]  
\[ -6.9561 = T_F - 65.5^\circ \text{C} \]  
\[ 58.5^\circ \text{C} = T_F \]
28. \[ m \text{CAT} = -m \text{CAT} \]
\[ (15.0 \text{g}) (4.184 \text{J/g} \cdot \text{C}) (50.0^\circ \text{C} - 30.0^\circ \text{C}) = -83.8 \text{J} \cdot \text{C} \cdot (50.0^\circ \text{C} - 60.0^\circ \text{C}) \]
\[ 0.276 \text{J} = 46090 \text{g} \cdot \text{C} \]
\[ 0.136 \text{J/g} \cdot \text{C} = \text{C}_{\text{metal}} \text{, tungsten} \]

33. \[ \text{C}_2\text{H}_5\text{OH}(s) + \text{Heat} \rightarrow \text{C}_2\text{H}_5\text{OH}(l) \]
\[ \frac{4.60 \text{g} \text{C}_2\text{H}_5\text{OH}}{1 \text{ mol} \text{C}_2\text{H}_5\text{OH}} \rightarrow 4.94 \text{ KJ} = 0.493 \text{ KJ} \]
\[ \frac{46.08 \text{g} \text{C}_2\text{H}_5\text{OH}}{1 \text{ mol} \text{C}_2\text{H}_5\text{OH}} \]

heat + \text{NH}_3(g) \rightarrow \text{NH}_3(g)

34. \[ 2450 \text{ KJ} \text{NH}_3 \rightarrow 1 \text{ mol} \text{NH}_3 \rightarrow 1790 \text{ gNH}_3 \]
\[ 23.3 \text{ KJ} \rightarrow 1 \text{ mol} \text{NH}_3 \]

heat + \text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(g)

35. \[ 63.7 \text{ g} \text{H}_2\text{O}(l) \rightarrow 1 \text{ mol} \text{H}_2\text{O} \rightarrow 40.7 \text{ KJ} = 144 \text{ KJ} \]
\[ \frac{18.02 \text{ g} \text{H}_2\text{O}}{1 \text{ mol} \text{H}_2\text{O}} \]
**Bond Energy**

What makes a reaction endothermic or exothermic?

**Why?**

Chemical reactions can be a lot like playing with Legos®—you must take apart part of your last creation before you can replace it with something new. For many chemical reactions, we have to first break bonds in the reactants before we can put the atoms back into a new arrangement to form the products. Both of these processes involve changes in energy. The net energy change for a reaction is called the **heat of reaction** or the **change in enthalpy** (ΔH). In this activity we will look at one way energy changes can be approximated for chemical reactions.

### Model 1 – Breaking and Forming Bonds

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Change in Energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>A) PCl₃(g) → P(g) + 3Cl(g)</td>
<td>+ 966.7 broken</td>
</tr>
<tr>
<td>B) PCl₃(g) → P(g) + 5Cl(g)</td>
<td>+ 1,297.9 broken</td>
</tr>
<tr>
<td>C) PF₅(g) → P(g) + 3F(g)</td>
<td>+ 1,470.4 broken</td>
</tr>
<tr>
<td>D) PF₅(g) → P(g) + 5F(g)</td>
<td>+ 2,305.4 broken</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Change in Energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>E) P(g) + 3Cl(g) → PCl₃(g)</td>
<td>− 966.7 formed</td>
</tr>
<tr>
<td>F) P(g) + 5Cl(g) → PCl₃(g)</td>
<td>− 1,297.9 formed</td>
</tr>
<tr>
<td>G) P(g) + 3F(g) → PF₅(g)</td>
<td>− 1,470.4 formed</td>
</tr>
<tr>
<td>H) P(g) + 5F(g) → PF₅(g)</td>
<td>− 2,305.4 formed</td>
</tr>
</tbody>
</table>

1. Identify four reactions in Model 1 where bonds are being broken. Write “broken” in the Change in Energy box in the table for each reaction you identify.

2. Identify four reactions in Model 1 where bonds are being formed. Write “formed” in the Change in Energy box in the table for each reaction you identify.

3. Circle the correct word to complete each sentence below.
   a. When bonds are **broken** there is a positive energy change.
   b. Breaking bonds is **endothermic** (exothermic).
   c. When bonds are **formed** there is a negative energy change.
   d. Forming bonds is **endothermic** (exothermic).

4. Find two reactions in Model 1 that are exact opposites of each other, that is, one reaction is the reverse of the other reaction.
   a. How do the changes in energy for the reverse reactions compare?
      *The changes in energy for these reactions are the same, but opposite in sign.*
   b. Explain your answer to part a considering what you learned from Questions 1–3 about bond breaking and bond formation.
      *If a given amount of energy is required to break a bond, then the same amount of energy will be released when the bond is formed.*
5. Consider the data in Model 1.

a. What unit is on the energy changes?

\[ \text{kJ/mol} \]

b. For Reaction A of Model 1, how many P–Cl bonds are broken with the 966.7 kJ of energy listed? *Hint: Look at the units you listed in part a.*

The energy change is listed per mole of reaction, corresponding therefore to breaking the P–Cl bonds in one mole of PCl₅. Three moles of P–Cl bond are broken.

c. Calculate the energy needed to break one mole of P–Cl bonds in Reaction A.

\[ \frac{966.7 \text{ kJ}}{1 \text{ mole PCl}_5} \times \frac{1 \text{ mole PCl}_5}{3 \text{ mole P–Cl}} = 322.2 \text{ kJ/mole} \]

6. Use the data in Model 1 to answer the following questions.

a. Calculate the energy needed to break one mole of P–Cl bonds in Reaction B.

\[ \frac{1,297.9 \text{ kJ}}{1 \text{ mole PCl}_5} \times \frac{1 \text{ mole PCl}_5}{5 \text{ mole P–Cl}} = 259.6 \text{ kJ/mole} \]

b. Do the P–Cl bonds in different molecules require the same amount of energy to break?

No.

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**Read This!**

The energies you calculated in Questions 5c and 6a above are called bond energies. The bond energy for a particular type of bond can vary from one molecule to another because the atomic environment of a bond can influence the amount of energy needed to break the bond. For example, the carbon–carbon bond in the two molecules shown below may not have the same bond energy because the surrounding atoms are different.

\[ \text{H—C—C—H} \quad \text{H—C—C—Cl} \]

Bond energies can be very useful (as you will soon discover) for calculating the net energy change in a reaction. However, a table listing the bond energies for even the most common substances would be several pages long. For this reason, chemists often approximate energy changes using average bond energy.
## Model 2 – Average Bond Energy

<table>
<thead>
<tr>
<th>Bond</th>
<th>Energy (kJ/mol)</th>
<th>Bond</th>
<th>Energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H—H</td>
<td>432</td>
<td>N—N</td>
<td>160</td>
</tr>
<tr>
<td>C—H</td>
<td>411</td>
<td>N=O</td>
<td>631</td>
</tr>
<tr>
<td>N—H</td>
<td>386</td>
<td>N≡N</td>
<td>941</td>
</tr>
<tr>
<td>H—Cl</td>
<td>431</td>
<td>N—O</td>
<td>201</td>
</tr>
<tr>
<td>C—C</td>
<td>346</td>
<td>Cl—Cl</td>
<td>243</td>
</tr>
<tr>
<td>C—O</td>
<td>358</td>
<td>F—F</td>
<td>158</td>
</tr>
<tr>
<td>C—N</td>
<td>305</td>
<td>O—H</td>
<td>464</td>
</tr>
<tr>
<td>C—Cl</td>
<td>327</td>
<td>O—Cl</td>
<td>269</td>
</tr>
<tr>
<td>C≡C</td>
<td>602</td>
<td>O—O</td>
<td>204</td>
</tr>
<tr>
<td>C≡O</td>
<td>745</td>
<td>C—F</td>
<td>552</td>
</tr>
<tr>
<td>O=O</td>
<td>494</td>
<td>C—S</td>
<td>259</td>
</tr>
</tbody>
</table>

7. Look at the bond energies in Model 2.
   a. Are the bond energies positive or negative?
   *The bond energies are listed as positive numbers.*
   b. Do the bond energies refer to bond-breaking or bond-forming processes?
   *Bond breaking.*

8. Look at the energy values in Model 2.
   a. What are the units for bond energy?
   *kJ/mol*
   b. How many bonds will the energy listed in the table break or form?
   *The values correspond to the energy required to break a mole of bonds, \(6.022 \times 10^{23}\) bonds.*

9. Is the bond energy for a double bond simply two times that of a single bond for the same atoms? Provide two specific examples to support your answer.
   *The double bond energy is a bit less than twice the single bond energy. Examples include C—O (358 kJ/mol) versus C≡O (745 kJ/mol).*

10. Explain how to use the information in Model 2 to determine the change in energy when a mole of C—H bonds is formed rather than broken.
    *Bond formation releases energy. The change in energy for forming C—H bonds will have the same numerical value as the bond energy listed in Model 2 but will be a negative value. Change the sign on the energy to negative.*

11. Because the environment surrounding a bond in a molecule alters the bond energy slightly, predict how scientists might calculate one bond energy value for the table in Model 2.
    *Take an average of all the known values for that type of bond.*
12. Write the overall reaction for the process that is illustrated in Model 3.

\[ N_2(g) + 3H_2(g) \rightarrow 2NH_3(g) \]

13. What does the y-axis on the graph in Model 3 represent?

*Change in potential energy*

14. Consider the process illustrated in Model 3.

a. Does Step 1 represent a bond-breaking or bond-forming process?

*Bond-breaking.*

b. Explain why the Step 1 arrow goes up—toward higher potential energy—rather than down.

*Breaking bonds requires that energy be absorbed, so the PE increases.*

15. Draw a Lewis dot diagram for *every* reactant molecule in Model 3.

\[ :\text{N} \equiv \text{N} : + \text{H} \equiv \text{H} \]

\[ \text{H} \equiv \text{H} \]

\[ \text{H} \equiv \text{H} \]

16. Use the bond energies in Model 2 to calculate the energy needed to break apart the four moles of molecules in Step 1 of the reaction in Model 3. Include the proper sign and units on your answer.

\[ 1(\text{N} \equiv \text{N \ bond}) + 3(\text{H} \equiv \text{H \ bond}) \]

\[ 1(941 \text{ kJ/mol}) + 3(432 \text{ kJ/mol}) = +2237 \text{ kJ/mol} \]

17. Draw a Lewis dot diagram for *every* product molecule in Model 3.

\[ \text{H} \equiv \text{N} \equiv \text{H} \quad \text{H} \equiv \text{N} \equiv \text{H} \]

\[ \text{H} \quad \text{H} \]
18. Use the bond energies in Model 2 to calculate the energy that is released to form the two moles of molecules in Step 2 of the reaction in Model 3. Include the proper sign and units on your answer.

\[ 6(N-H \text{ bond}) \]
\[ 6(-386 \text{ kJ/mol}) = -2,316 \text{ kJ/mol} \]

*Energy is released when bonds are formed. Use negative number for the bond energy.*

19. The heat of reaction or change in enthalpy (\(\Delta H\)) of a reaction is the net energy change in the reaction. Consider where \(\Delta H\) is located in Model 3. Discuss with your group members how your answers to Questions 16 and 18 could be used to calculate the enthalpy change (\(\Delta H\)) for the reaction in Model 3, and then do the calculation.

*Add the energy absorbed (positive) to the energy released (negative) to calculate \(\Delta H\).*

\[ 2,237 \text{ kJ/mol} + (-2,316 \text{ kJ/mol}) = -79 \text{ kJ/mol} \]

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**Read This!**

The diagram in Model 3 suggests that the production of ammonia from its elements occurs by breaking the element molecules into single gaseous atoms and then recombining them into ammonia. In actuality, very few chemical reactions occur in this manner. In fact, the production of ammonia does not happen this way. Lucky for chemists, it doesn’t really matter. Enthalpy is a state function, which means the actual process or pathway by which the reaction takes place does not affect the net energy change. Therefore, it is perfectly alright to imagine that the reaction occurs as illustrated in Model 3. The enthalpy change that is calculated will still approximate the real-life reaction because for a state function like enthalpy only the beginning state (reactants) and the ending state (products) must be known.

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20. Balance the following reaction.

\[ \text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \]
\[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \]


\[ \begin{array}{c}
\text{H} \\
\text{H-\text{C-H}} \\
\text{H}
\end{array} \quad :\ddots:\ddots: \quad \begin{array}{c}
\text{H} \\
\text{H-\text{\ddots-H}} \\
\text{H}
\end{array} \quad \begin{array}{c}
:\ddots: \\
\text{\ddots-O-C-\ddots:}
\end{array} \quad \begin{array}{c}
\text{H-\ddots-H} \\
\text{H-\ddots-H}
\end{array} \]

22. Use the average bond energy in Model 2 to calculate the change in enthalpy (\(\Delta H\)) for the reaction in Question 20.

\[ \begin{array}{c}
\text{broken} \\
4(\text{C-H bonds}) + 2(\text{O=O bonds}) + 2(\text{C=O bonds}) + 4(\text{H-O bonds})
\end{array} \quad \begin{array}{c}
\text{formed} \\
4(411 \text{ kJ/mol}) + 2(494 \text{ kJ/mol}) + 2(-799 \text{ kJ/mol}) + 4(-464 \text{ kJ/mol}) = -822 \text{ kJ/mol}
\end{array} \]

23. Why are the changes in enthalpy calculated using average bond energies considered approximate values?

*The values in the bond energy table in Model 2 are average bond energies, so they are not the actual bond energies for the particular molecules used in the reaction. Therefore, any change in enthalpy that is calculated will be approximate. For example, the experimental enthalpy for \( \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \) is \(-92.2 \text{ kJ/mol}\). A bit different than the answer calculated in Question 19, but it is close.*

Bond Energy
Extension Questions

Model 4 – C—C Bond versus C=C Bond

24. The graphs in Model 4 show the change in potential energy as a function of the distance between two atoms as they approach each other to form a bond. What two bonds do the graphs represent?

\[ \text{C—C and C=C} \]

25. Consider the information in Model 4.

\( a. \) What happens to the potential energy of the atoms as they get closer to each other and form a bond?

\[ \text{The PE decreases.} \]

\( b. \) Explain your answer in part \( a. \) Your answer should include a discussion of attractive and repulsive forces.

\[ \text{As the bond forms, the attractive forces between the two nuclei and the electrons are increasing more than the repulsive forces between nuclei or between electron clouds.} \]

\( c. \) Does this difference in potential energy cause a release of energy or absorption of energy from the atoms?

\[ \text{The potential energy decreases, resulting in a release of energy as the bond forms.} \]

26. Consider the information in Model 4.

\( a. \) What happens to the potential energy of the atoms if they get too close to each other?

\[ \text{The PE increases.} \]

\( b. \) Explain your answer in part \( a. \) Your answer should include a discussion of attractive and repulsive forces.

\[ \text{As the atoms get too close, the repulsive force between the two nuclei increases more than the attractive forces between the nuclei and electron clouds. Also, the electron clouds begin to repel each other more at closer distances.} \]
27. Compare the two graphs in Model 4.

a. Which bond in Model 4 (the single bond or double bond) will release a larger amount of potential energy when it forms?

\[ \text{C}=\text{C} \]

b. Which bond in Model 4 (the single bond or double bond) will have a larger average bond enthalpy value?

\[ \text{C}=\text{C} \]

c. Which bond in Model 4 (the single bond or double bond) has a shorter bond length?

\[ \text{C}=\text{C} \]

d. Explain how the graphs in Model 4 help you answer the question in part c.

*The x-axis of each graph shows distance. The lowest point on the y-axis, corresponding to the minimum potential energy, for the graph on the left occurs further to the right on the x-axis scale, corresponding to a longer distance than the graph on the right. That means the double bond is more stable at a shorter distance than the single bond.*
Heats of Formation

What do the tabulated $\Delta H_f$ values in the appendix of my textbook mean?

Why?

Enthalpies of formation, $\Delta H_f$, are commonly tabulated in chemistry texts and reference materials. They are the most common way of calculating theoretical $\Delta H^\circ$ values via Hess's law, but often the reaction that they refer to is not shown in the table. In this activity, we will explore these values, what they represent and how they can be used to predict the enthalpies of other reactions.

Model 1 – Comparing Reactions

Set A

<table>
<thead>
<tr>
<th>Formation Reactions</th>
<th>$\Delta H_f$ (kJ/mole)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\frac{3}{2}O_2(\text{g}) + 2\text{Al}(\text{s}) \to Al_2O_3(\text{s})$</td>
<td>-1675</td>
</tr>
<tr>
<td>$6\text{C}(\text{graphite}) + 6\text{H}_2(\text{g}) + 3\text{O}<em>2(\text{g}) \to C_6\text{H}</em>{12}\text{O}_6(\text{s})$</td>
<td>-6984</td>
</tr>
<tr>
<td>$\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \to \text{H}_2\text{O}(\text{l})$</td>
<td>-286</td>
</tr>
<tr>
<td>$\frac{1}{2}\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \to \text{NO}_2(\text{g})$</td>
<td>+33</td>
</tr>
<tr>
<td>$\frac{1}{2}\text{H}_2(\text{g}) + \text{C}(\text{graphite}) + \frac{1}{2}\text{N}_2(\text{g}) \to \text{HCN}(\text{g})$</td>
<td>+131</td>
</tr>
</tbody>
</table>

Set B

<table>
<thead>
<tr>
<th>Not Formation Reactions</th>
<th>$\Delta H^\circ$ (kJ/mole)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$3\text{O}_2(\text{g}) + 4\text{Al}(\text{s}) \to 2\text{Al}_2\text{O}_3(\text{s})$</td>
<td>-3350</td>
</tr>
<tr>
<td>$\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \to \text{C}<em>6\text{H}</em>{12}\text{O}_6(\text{s})$</td>
<td>+2805</td>
</tr>
<tr>
<td>$2\text{H}(\text{g}) + \text{O}(\text{g}) \to \text{H}_2\text{O}(\text{l})$</td>
<td>-1121</td>
</tr>
<tr>
<td>$\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \to \text{H}_2\text{O}(\text{l})$</td>
<td>-231</td>
</tr>
<tr>
<td>$2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \to 2\text{NO}_2(\text{g})$</td>
<td>-114</td>
</tr>
</tbody>
</table>

1. Consider the reactions in Model 1. What type of reaction are all of the examples in Model 1—synthesis, decomposition, single replacement, double replacement or combustion?

   *Synthesis.*

2. Compare the first reaction in each set (producing aluminum oxide).

   a. How are the reactions similar?

   *Both reactions produce aluminum oxide. Both reactions use oxygen gas and solid aluminum as reactants.*

   b. How are the reactions different?

   *The coefficients for the reaction in Set B are double the coefficients for the reaction in Set A.*
3. Compare the two sets of reactions in Model 1.
   a. Which set of reactions uses fractional coefficients to balance the equation, forcing the product to always have a coefficient of one?
      
      *Some reactions in Set A use fractions to balance the equation.*
   b. Which set of reactions uses only elements as reactants?
      
      *Set A uses only elements as reactants.*
   c. Which set of reactions uses reactants that would not be stable substances (not their natural state)? For example, a single oxygen atom is not a stable entity.
      
      *Some reactions in Set B use reactants that would not be in a stable state.*

4. A **formation reaction** is defined as a synthesis reaction where one mole of product is made from component elements in their natural state.
   a. Which set of reactions in Model 1 are formation reactions?
      
      *Set A.*
   b. Add a label to Model 1 indicating this set as a group of formation reactions.
      
      *See Model 1.*

5. Explain why the following reactions are **not** formation reactions.
   
   **Mg^{2+} (aq) + 2Cl^- (aq) → MgCl_2 (s)** *Reactants are ions, not elements in their natural state.*
   
   **2Na(s) + 2H_2O (l) → 2NaOH (aq) + H_2 (g)** *Reactants contain a compound.*
   
   **2H_2 (g) + O_2 (g) → 2H_2O (l)** *Reaction is not balanced to form 1 mole of product.*

6. The numerical values in Model 1 are standard enthalpies for the reactions. Enthalpies for formation reactions are called standard heats of formation, and they are given the symbol ΔH° f.
   a. Label the column of enthalpies for the formation reactions ΔH° f.
      
      *See Model 1.*
   b. Label the column of enthalpies for the other reactions ΔH°.
      
      *See Model 1.*

7. Are standard heats of formation endothermic, exothermic or both?
   
   *Model 1 shows that some standard heats of formation are endothermic (positive) while others are exothermic (negative).*

8. Consider what you know about the energy involved in breaking and forming bonds. Explain why some standard heats of formation might be endothermic?
   
   *Some formation reactions require the breaking of bonds in reactant molecules, e.g., O_2 (g) or Al(s). The energy required to break these bonds may be larger than the energy released when the products are formed, making the overall reaction endothermic.*
9. Find a table of standard heats of formation in your textbook or online.
   a. What is the $\Delta H^\circ_f$ for solid sodium chloride?
      \[-411 \text{ kJ/mole}\]
   
   b. Write the reaction that involves that energy change.
      \[\text{Na}(s) + \frac{1}{2}\text{Cl}_2(g) \rightarrow \text{NaCl}(s)\]
   
   c. What is the $\Delta H^\circ_f$ for gaseous iodine, $\text{I}_2(g)$?
      \[62.4 \text{ kJ/mole}\]
   
   d. Write the reaction that involves that energy change.
      \[\text{I}_2(s) \rightarrow \text{I}_2(g)\]

10. Explain why the $\Delta H^\circ_f$ for gaseous oxygen, $\text{O}_2(g)$, is zero.

   *Oxygen gas is already an element in its natural state, so no energy is needed to produce it.*

11. If a formation reaction was reversed, how could the standard enthalpy for that reaction be calculated from a standard heat of formation listed in the table? Illustrate your understanding by finding the $\Delta H^\circ$ for the following reaction.

   \[\text{HBr}(g) \rightarrow \frac{1}{2}\text{H}_2(g) + \frac{1}{2}\text{Br}_2(l)\]

   *When a reaction is reversed, the $\Delta H$ of the reaction switches sign from positive to negative or vice versa. The $\Delta H^\circ_f$ of hydrogen bromide is $-36$ kJ/mole, so the $\Delta H^\circ$ for this reaction would be $+36$ kJ/mole.*

12. If a formation reaction was multiplied by a factor, how could the standard enthalpy for that reaction be calculated from a Standard Heat of Formation listed in the table? Illustrate your understanding by finding the $\Delta H^\circ$ for the following reaction.

   \[2\text{Al}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{AlCl}_3(s)\]

   *When a reaction is multiplied by a factor, the $\Delta H$ of the reaction is also multiplied by that factor. The $\Delta H^\circ$ of aluminum chloride is $-706$ kJ/mole, so the $\Delta H^\circ$ for this reaction would be $-1412$ kJ/mole.*

Read This!

Standard heats of formation can be used to predict the enthalpies of many other reactions. Imagine that a reaction occurs by all reactants decomposing into elements in their natural state and then recombining to form the products. Although very few chemical reactions occur in this manner, it does not really matter. Enthalpy is a *state function*, which means it is determined only by the initial state of the reaction and the final state of the reaction, not the process that occurs in between. The enthalpy change that is calculated for our imaginary process will give us the enthalpy for the real-life reaction. Therefore, several formation reactions (or reverse formation reactions) can be combined using Hess’s law to calculate the enthalpy for any reaction.
**Model 2 – Calculating an Enthalpy, \( \Delta H^\circ \), from \( \Delta H^\circ_f \) Values**

<table>
<thead>
<tr>
<th>Potential Energy</th>
<th>Intermediates</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>C(graphite)(s) + 2H(_2)(g) + 2O(_2)(g)</td>
</tr>
<tr>
<td>Step 1</td>
<td>Step 2</td>
</tr>
<tr>
<td>CH(_4)(g) + 2O(_2)(g)</td>
<td>CO(_2)(g) + 2H(_2)O(g)</td>
</tr>
<tr>
<td>Step 3</td>
<td>Step 4</td>
</tr>
</tbody>
</table>

### 13. Write the reaction for the overall process in Model 2.

\[
CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)
\]

### 14. Notice that the intermediates in the process in Model 2 are elements in their natural states.

**a.** Which steps in the process in Model 2 represent formation reactions, or multiples of formation reactions?

*Steps 3 and 4 are formation reactions.*

**b.** Which steps in the process in Model 2 represent formation reactions, or multiples of formation reactions, in reverse?

*Steps 1 and 2 are reverse formation reactions.*

### 15. Use a table of standard heats of formation to calculate the energy change when the reactants in Model 2 turn into the intermediates. Be sure to include the correct sign on this energy change.

- \( CH_4(g) \quad \Delta H^\circ_f = -75 \text{ kJ/mole} \)
- \( O_2(g) \quad \Delta H^\circ_f = 0 \text{ kJ/mole} \)
- \( Step 1 + Step 2 = +75 \text{ kJ/mole} + 2(0 \text{ kJ/mole}) = +75 \text{ kJ/mole} \)

### 16. Use a table of standard heats of formation to calculate the energy change when the intermediates in Model 2 turn into the products. Be sure to include the correct sign on this energy change.

- \( CO_2(g) \quad \Delta H^\circ_f = -394 \text{ kJ/mole} \)
- \( H_2O(g) \quad \Delta H^\circ_f = -242 \text{ kJ/mole} \)
- \( Step 3 + Step 4 = -394 \text{ kJ/mole} + 2 \text{ moles}(-242 \text{ kJ/mole}) = -878 \text{ kJ/mole} \)

### 17. Use your answers in Questions 15 and 16 to calculate the overall \( \Delta H^\circ \) for the process in Model 2.

\[
75 \text{ kJ/mole} + (-878 \text{ kJ/mole}) = -803 \text{ kJ/mole}
\]

### 18. Use what you have learned from Model 2 to calculate the standard enthalpy, \( \Delta H^\circ \), for the following reaction using a table of standard heats of formation. Show all calculations and be prepared to explain your process for solving this problem.

\[ 4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g) \]

\[
\Delta H^\circ_f = -46 \text{ kJ/mole} \quad 0 \text{ kJ/mole} \quad 90 \text{ kJ/mole} \quad -242 \text{ kJ/mole}
\]

\[
\Delta H = 4 \text{ moles}(-46 \text{ kJ/mole}) + 5 \text{ moles}(0 \text{ kJ/mole}) + 4 \text{ moles}(90 \text{ kJ/mole}) + 6 \text{ moles}(-242 \text{ kJ/mole}) = -908 \text{ kJ}
\]
**Extension Questions**

19. The heat of combustion for 1 mole of propane, $C_3H_8$, at standard conditions is $-2,046$ kJ.
   
   a. Write a balanced chemical equation for the combustion of propane.
   
   \[ C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g) \]
   
   b. Calculate the heat of formation, $\Delta H_f^\circ$, of propane given that the $\Delta H_f^\circ$ of gaseous water is $-242$ kJ/mole and the $\Delta H_f^\circ$ of carbon dioxide gas is $-394$ kJ/mole.
   
   \[ \Delta H_f^\circ_{\text{comb}} = 1 \text{ mole}(\Delta H_f^\circ_{\text{propane}}) + 5 \text{ moles}(\Delta H_f^\circ_{\text{oxygen}}) + 3 \text{ moles}(\Delta H_f^\circ_{\text{carbon dioxide}}) + 4 \text{ moles}(\Delta H_f^\circ_{\text{gaseous water}}) \]
   
   \[ = -2,046 \text{ kJ} = 1 \text{ mole}(\Delta H_f^\circ_{\text{propane}}) + 5 \text{ moles}(0 \text{ kJ/mole}) + 3 \text{ moles}(-394 \text{ kJ/mole}) + 4 \text{ moles}(-242 \text{ kJ/mole}) \]
   
   \[ \Delta H_f^\circ_{\text{propane}} = -104 \text{ kJ/mole} \]

20. Which of the two formation reactions below would you predict to have the largest exothermic heat of formation? Justify your reasoning.

   \[ 2\text{Al}(s) + O_2(g) \rightarrow \text{Al}_2\text{O}_3(s) \]

   \[ \Delta H_f^\circ = -1675 \text{ kJ/mole} \]

   \[ \text{Mg}(s) + \frac{1}{2}O_2(g) \rightarrow \text{MgO}(s) \]

   \[ \Delta H_f^\circ = -601 \text{ kJ/mole} \]

*The heat of formation for aluminum oxide is larger because the Al$_2$O$_3$ crystal lattice is stronger than the MgO crystal lattice. Aluminum has a +3 charge while magnesium only has a +2 charge. The Coulombic attraction between aluminum and oxygen is stronger. The energy needed to break the bonds on the reactant side would be similar—separating the metal atoms in the solid and breaking bonds in the gaseous oxygen atoms.*
Honors Hess’s Law Worksheet

Find the ΔH for the underlined reaction, given the following reactions and subsequent ΔH values. Also state whether the reaction is endothermic or exothermic.

1. \(2\text{C}_2\text{H}_4\text{O}(l) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{C}_2\text{H}_5\text{O}(l) + \text{O}_2(g)\)
   \[\Delta H = -204 \text{kJ, endo}\]

2. \(\text{RC}_2\text{H}_4\text{O}(l) + 1\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(l)\)
   \[\Delta H = -685.5 \text{kJ}\]

   \(2\text{RC}_2\text{H}_4\text{O}(l) + 5\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 4\text{H}_2\text{O}(l)\)
   \[\Delta H = -1167 \text{kJ}\]

   \(\Delta H = -1628.2 \text{kJ, exo}\)

3. \(4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g)\)
   \[\Delta H = -1628.2 \text{kJ, exo}\]

   \(\text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g)\)
   \[\Delta H = -180.5 \text{kJ}\]

   \(\text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)\)
   \[\Delta H = -91.8 \text{kJ}\]

   \(2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g)\)
   \[\Delta H = -483.6 \text{kJ}\]

   \(\Delta H = -486 \text{kJ, exo}\)

4. \(\text{H}_2\text{SO}_4(l) \rightarrow \text{SO}_2(g) + \text{H}_2\text{O}(g)\)
   \[\Delta H = -73 \text{kJ, endo}\]

   \(\text{H}_2\text{S}(g) + 2\text{O}_2(g) \rightarrow \text{H}_2\text{SO}_4(l)\)
   \[\Delta H = -235.5 \text{kJ}\]

   \(\Delta H = -207 \text{kJ}\)

   \(\Delta H = 44 \text{kJ}\)

   \(\Delta H = 72.5\)

5. \(\text{N}_2\text{H}_4(l) + \text{H}_2(g) \rightarrow 2\text{NH}_3(g)\)
   \[\Delta H = -18 \text{kJ, exo}\]

   \(\text{N}_2\text{H}_4(l) + \text{CH}_4\text{O}(l) \rightarrow \text{CH}_2\text{O}(g) + \text{N}_2(g) + 3\text{H}_2(g)\)
   \[\Delta H = -37 \text{kJ}\]

   \(\Delta H = -46 \text{kJ}\)

   \(\text{R}_\text{CH}_4\text{O}(l) \rightarrow \text{CH}_2\text{O}(g) + \text{H}_2(g)\)
   \[\Delta H = -65 \text{kJ}\]

   \[\Delta H = -18\]
For question 6-10, the some thermochemical equations need to be multiplied by a fraction.

6. \[ \text{CH}_4 (g) + \text{NH}_3 (g) \rightarrow \text{HCN} (g) + 3 \text{H}_2 (g) \quad \Delta H = 256.0 \text{kJ} \text{ Endo} \]
   \[ \frac{1}{2} \text{N}_2 (g) + 3 \text{H}_2 (g) \rightarrow 2 \text{NH}_3 (g) \quad \Delta H = -91.8 \text{kJ} \]
   \[ \text{H}_2 (g) + 2 \text{C} (s) + \frac{1}{2} \text{N}_2 (g) \rightarrow 2 \text{HCN} (g) \quad \Delta H = +270.3 \text{kJ} \]
   \[ \text{Endo} = \frac{255.95 - 256.0}{2} = 0.0 \text{kJ} \]

7. \[ \text{NO}_2 (g) + \frac{7}{2} \text{H}_2 (g) \rightarrow 2 \text{H}_2 \text{O}(l) + \text{NH}_3 (g) \quad \Delta H = -108 \text{kJ} \text{ Endo} \]
   \[ \frac{1}{2} \text{N}_2 (g) + 3 \text{H}_2 (g) \rightarrow 2 \text{NH}_3 (g) \quad \Delta H = -46 \text{kJ} \]
   \[ \text{NO}_2 (g) + 2 \text{H}_2 (g) \rightarrow \text{H}_2 \text{O}(l) \quad \Delta H = -85 \text{kJ} \]
   \[ \sum \Delta H = -23 - 85 = -108 \text{kJ} \approx 250 \text{kJ} \text{ Endo} \]
   \[ \text{or} \quad 250 \cdot \text{kJ} = 2.50 \times 10^2 \text{kJ} \text{ Endo} \]

8. \[ \text{PCl}_3 (g) \rightarrow \text{PCl}_3 (g) + \text{Cl}_2 (g) \quad \Delta H = ? \]
   \[ \frac{1}{4} \text{P}_4 (s) + 6 \text{Cl}_2 (g) \rightarrow 4 \text{PCl}_3 (g) \quad \Delta H = -609.75 \text{kJ} \]
   \[ \frac{1}{4} \text{P}_4 (s) \rightarrow 10 \text{Cl}_2 (g) \quad \Delta H = 3438 \text{kJ} \]
   \[ \frac{1}{4} \text{P}_4 (s) \rightarrow 10 \text{Cl}_2 (g) \quad \Delta H = 3438 \text{kJ} \]
   \[ \sum \Delta H = 249.75 \approx 2.50 \times 10^2 \text{kJ} \] 

9. \[ 2 \text{CO}_2 (g) + \text{H}_2 \text{O}(g) \rightarrow \text{C}_2 \text{H}_2 (g) + 5/2 \text{O}_2 (g) \quad \Delta H = ? \]
   \[ \text{C}_2 \text{H}_2 (g) + 2 \text{H}_2 (g) \rightarrow \text{C}_2 \text{H}_6 (g) \quad \Delta H = -94.5 \text{kJ} \]
   \[ \text{H}_2 \text{O}(g) \rightarrow 2 \text{H}_2 (g) + \text{O}_2 (g) \quad \Delta H = 142.4 \text{kJ} \]
   \[ \text{C}_2 \text{H}_6 (g) + 7/2 \text{O}_2 (g) \rightarrow 4 \text{CO}_2 (g) + 6 \text{H}_2 \text{O}(g) \quad \Delta H = -566 \text{kJ} \]
   \[ \sum \Delta H = 94.5 + 142.4 + 2.83 = 235 \text{kJ} \text{ Endo} \]

Most Challenging!!!

10. \[ \text{FeO(s)} + \text{CO(g)} \rightarrow \text{Fe(s)} + \text{CO}_2 (g) \quad \Delta H = ? \]
    \[ \frac{1}{6} \text{Fe}_2 \text{O}_3 (s) + \text{CO(g)} \rightarrow 2 \text{Fe}_2 \text{O}_4 (s) + \text{CO}_2 (g) \quad \Delta H = -47.0 \text{kJ} \]
    \[ \frac{1}{3} \text{Fe}_2 \text{O}_4 (s) + 3 \text{CO(g)} \rightarrow 2 \text{Fe(s)} + 3 \text{CO}_2 (g) \quad \Delta H = -25.0 \text{kJ} \]
    \[ \frac{1}{6} \text{Fe}_3 \text{O}_4 (s) + \text{CO(g)} \rightarrow 3 \text{FeO(s)} + \text{CO}_2 (g) \quad \Delta H = 19.0 \text{kJ} \]
    \[ \sum \Delta H = 7.83 + 12.5 + 19.0 = 40.3 \text{kJ} \text{ Endo} \]
    \[ \text{or} \quad 40.3 \times \text{kJ} = 40.3 \times 10^2 \text{kJ} \text{ Endo} \]
Calorimetry

What is the relationship between heat energy and temperature?

Why?

When a substance is heated, the temperature of that substance increases. Will the same amount of energy cause different substances to have identical temperature increases? Will the same amount of energy be needed to cause identical temperature increases in different amounts of the same substance? In this activity you will explore how mass, temperature, heat energy, and the type of substance are related.

Model 1 – A Pot of Water

Before heating

![4 qt. saucepan temperature 24.5°C](image1)

![8 qt. stockpot temperature 24.5°C](image2)

After heating for 5 minutes at a rate of 30 Joules per minute

![4 qt. saucepan temperature](image3)

![8 qt. stockpot temperature](image4)

1. In Model 1, which container holds more grams of water?
   
The 8 qt. stockpot holds more grams of water.

2. Consider the process described in Model 1:
   
a. How many joules of energy were added to the saucepan?
   
The heating rate was 30 joules per minute. A total of 150 J was added to the saucepan in 5 minutes.

b. How many joules of energy were added to the stockpot?
   
The heating rate was 30 joules per minute. A total of 150 J was added to the stockpot in 5 minutes.

c. In which container did the liquid gain more energy or did both gain the same amount? Explain your reasoning.
   
Both liquids gained the same amount of energy because the heating rate was the same.

d. For each container, include whether the temperature is expected to increase, decrease or remain the same after heating. Explain your reasoning.
   
The temperature should increase in both pots. However, the water in the two pots will reach different temperatures. The 4 qt. pot will reach higher temperatures since it contains a smaller volume of water yet absorbs the same amount of heat energy.

Calorimetry
### Model 2 – Experimental Data for Heating Water

#### Experiment 1

<table>
<thead>
<tr>
<th>Trial</th>
<th>Mass (g)</th>
<th>ΔT (°C)</th>
<th>Added Energy (Joules)</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>1.00</td>
<td>26.8</td>
<td>112</td>
</tr>
<tr>
<td>B</td>
<td>2.00</td>
<td>13.4</td>
<td>112</td>
</tr>
<tr>
<td>C</td>
<td>2.50</td>
<td>10.7</td>
<td>112</td>
</tr>
<tr>
<td>D</td>
<td>3.00</td>
<td>8.93</td>
<td>112</td>
</tr>
<tr>
<td>E</td>
<td>3.50</td>
<td>7.66</td>
<td>112</td>
</tr>
<tr>
<td>F</td>
<td>4.00</td>
<td>6.70</td>
<td>112</td>
</tr>
</tbody>
</table>

#### Experiment 2

<table>
<thead>
<tr>
<th>Trial</th>
<th>Mass (g)</th>
<th>ΔT (°C)</th>
<th>Added Energy (Joules)</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>3.00</td>
<td>3.00</td>
<td>37.6</td>
</tr>
<tr>
<td>B</td>
<td>3.00</td>
<td>6.00</td>
<td>75.2</td>
</tr>
<tr>
<td>C</td>
<td>3.00</td>
<td>9.00</td>
<td>113</td>
</tr>
<tr>
<td>D</td>
<td>3.00</td>
<td>12.0</td>
<td>150</td>
</tr>
<tr>
<td>E</td>
<td>3.00</td>
<td>15.0</td>
<td>188</td>
</tr>
<tr>
<td>F</td>
<td>3.00</td>
<td>18.0</td>
<td>226</td>
</tr>
</tbody>
</table>

#### Experiment 3

<table>
<thead>
<tr>
<th>Trial</th>
<th>Mass (g)</th>
<th>ΔT (°C)</th>
<th>Added Energy (Joules)</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>1.00</td>
<td>9.00</td>
<td>37.6</td>
</tr>
<tr>
<td>B</td>
<td>2.00</td>
<td>9.00</td>
<td>75.0</td>
</tr>
<tr>
<td>C</td>
<td>2.50</td>
<td>9.00</td>
<td>94.1</td>
</tr>
<tr>
<td>D</td>
<td>3.00</td>
<td>9.00</td>
<td>113</td>
</tr>
<tr>
<td>E</td>
<td>3.50</td>
<td>9.00</td>
<td>132</td>
</tr>
<tr>
<td>F</td>
<td>4.00</td>
<td>8.97</td>
<td>150</td>
</tr>
</tbody>
</table>

3. In the data tables in Model 2, what does ΔT mean?
   *Change in temperature (temperature difference, final T – initial T).*

4. Which experiment in Model 2 illustrates how the amount of energy needed to achieve the same temperature change depends on the mass of water?
   *Experiment 3.*

5. Which experiment in Model 2 illustrates how different amounts of energy result in different temperature changes when the mass of water is constant?
   *Experiment 2.*

6. Why is it necessary to perform three experiments to find the relationships between mass, temperature change, and energy?
   *In order to determine the relationship between three different variables, you need to keep one constant in order to relate the other two. Therefore, three experiments are needed, keeping a different variable constant with each experiment.*
7. Refer to Experiment 1 in Model 2, and consider the relationship between the mass of water and the observed temperature change when the same amount of energy is added.

   a. Complete the statement below to show the relationship.

      When the same amount of energy is added to water samples of different mass, the change in temperature gets smaller/larger as the mass of the water increases.

   b. Does the relationship stated in part a describe a direct or inverse relationship?

      \textit{Inverse.}

8. Refer to Experiment 2 in Model 2.

   a. Write a grammatically correct sentence (like the one in Question 7a) to describe the relationship between the observed temperature change and the energy required to heat water samples of identical mass.

      \textit{When the same mass of water is heated, a larger temperature change requires more energy.}

   b. Does the relationship stated in part a describe a direct or inverse relationship?

      \textit{Direct.}

9. In Experiment 2, should the value of the missing energy in the last row be larger or smaller than the other energy values in that column?

      \textit{Larger.}

10. Refer to Experiment 3 in Model 2.

   a. Write a grammatically correct sentence to describe the relationship between the mass of water and the energy required to produce the same temperature change in different water samples.

      \textit{To get the same temperature change in two samples of water, the larger mass of water will require more energy.}

   b. Does the relationship stated in part a describe a direct or inverse relationship?

      \textit{Direct.}

11. Each of the experiments in Model 2 can be expressed as a proportionality indicated by the following symbol: \(\propto\). Several possible equations are shown in the table below.

\[
\begin{array}{c|c|c}
q & \propto & m \\
\hline
q & \propto & \Delta T \\
\Delta T & \propto & m \\
\end{array}
\]

   a. Describe what each variable represents in the equations shown above?

      \(m = \text{mass}\) \hspace{0.5cm} \(\Delta T = \text{change in temperature}\) \hspace{0.5cm} \(q = \text{energy}\)

   b. For each experiment in Model 2, indicate which of the six proportionalities shown in the table above correctly describes the relationship observed by the results of that experiment.

   Experiment 1: \(\Delta T \propto 1/m\) \hspace{1cm} Experiment 2: \(q \propto \Delta T\) \hspace{1cm} Experiment 3: \(q \propto m\)

\textbf{Calorimetry}
12. Based on your answers in Question 11, write a single proportionality statement that includes all three variables: q, m, ΔT. The statement should use q as the variable on the left.

\[ q \propto m \Delta T \]

13. A proportionality statement can become an equation when a constant is determined. The equation will have q on the left side of the equal sign and a constant “c” on the right side of the equation. Re-write the proportionality statement from Question 12 as an equation using “c” for the constant. After reaching consensus on this equation in your group, compare your answer with other groups to verify your equation.

\[ q = mc\Delta T \]

14. Solve for the value of the constant “c” using data from Model 2. Your teacher may assign you a specific experiment to use. Your group should calculate the value of c by using one set of data from each of the three experiments (include the units of this constant). After you complete the three calculations, compare your value for the constant with the value that other groups determined.

*Calculations will vary.*

**Sample calculation from Experiment 1, Trial C.**

\[ 112 \text{ J} = (2.50 \text{ g})(c)(10.7 \degree \text{C}) \]

\[ c = 4.18 \text{ J/g} \cdot \degree \text{C} \]

15. The constant obtained in Question 14 is called the **specific heat**, (c). It is an intensive physical property that has a different, characteristic value for every substance. What is the value for the specific heat of water?

\[ c = 4.18 \text{ J/g} \cdot \degree \text{C} \]

16. What do the units for specific heat mean? (Make sure your answer is a grammatically correct sentence.)

*You need 4.18 Joules of energy to increase the temperature of 1.0 gram of water by 1.0 degree Celsius.*

17. Using your equation for energy transferred (q) from Question 13 and the calculated value for the specific heat of water, determine the amount of energy required to increase the temperature of 550 grams of water by 20.0 °C.

\[ q = (550 \text{ g})(4.18 \text{ J/g} \cdot \degree \text{C})(20.0 \degree \text{C}) \]

\[ q = 46,000 \text{ J} \]

18. Using your equation for transferred energy (q) and the specific heat, determine the values for the missing data in the three experiments in Model 2. (Trial F in experiments 1–3.)

**See Model 2.**

19. Use your equation from Question 13 to calculate the following:

a. How much energy is transferred when 30.0 g of water is cooled from 25.0 °C to 12.7 °C.

\[ q = mc\Delta T \]

\[ q = (30.0 \text{ g})(4.18 \text{ J/g} \cdot \degree \text{C})(-12.3 \degree \text{C}) \]

\[ q = -1540 \text{ J} \text{ (rounded)} \]

b. Describe the significant difference between this value and the energy values shown in Model 2.

*The calculated value for q is a negative quantity because energy has been lost as the water cools (ΔT is negative).*

POGIL™ Activities for High School Chemistry
Extension Questions

Model 3 – Heating Mercury and Water

<table>
<thead>
<tr>
<th>Trial</th>
<th>Substance</th>
<th>Mass (grams)</th>
<th>ΔT (°C)</th>
<th>Added Energy (Joules)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Hg</td>
<td>100.0</td>
<td>71.0</td>
<td>1,000.0</td>
</tr>
<tr>
<td>2</td>
<td>Hg</td>
<td>100.0</td>
<td>142</td>
<td>2,000.0</td>
</tr>
<tr>
<td>3</td>
<td>H₂O</td>
<td>100.0</td>
<td>2.39</td>
<td>1,000.0</td>
</tr>
<tr>
<td>4</td>
<td>H₂O</td>
<td>100.0</td>
<td>4.78</td>
<td>2,000.0</td>
</tr>
</tbody>
</table>

20. What two substances are being heated in the experiments described in Model 3?

_Mercury and water._

21. Compare any two trials in Model 3 that use the same amount of energy but involve different substances. Is the temperature change the same or different in the two trials?

_Trials 1 and 3—Both samples had the same mass and amount of energy added. However, the mercury sample had a much larger temperature change._

22. Based on the data in this activity, the specific heat of water was determined to be 4.18 J/ g°C. Use data from Model 3 to solve for the value of the specific heat for mercury.

\[ q = mc\Delta T \]

\[ 1,000 \, J = (100 \, g) \, (c) \, (71.0 \, ^\circ C) \]

\[ c = 0.141 \, J/g\cdot ^\circ C \]

23. Select the correct word to complete the sentence:

When adding the same amount of energy to two similar massed samples, the substance with the larger specific heat will have a ___ temperature change?

(larger or smaller)

24. If 23,000 joules of energy are used to heat water by 4.00 °C, what is the mass of the water?

\[ 23,000 \, J = m \, (4.18 \, J/g\cdot ^\circ C) \, (4.00 \, ^\circ C) \]

\[ m = 1400 \, g \]

25. If 23,000 joules of energy are used to heat mercury by 4.00 °C, what is the mass of the mercury?

\[ 23,000 \, J = m \, (0.141 \, J/g\cdot ^\circ C) \, (4.00 \, ^\circ C) \]

\[ m = 41,000 \, g \]

26. What is the specific heat of aluminum if 4,750 joules of heat energy added to 249 g of aluminum produces a recorded temperature change of 21.1 °C?

\[ 4,750 \, J = (249 \, g) \, (c) \, (21.1 \, ^\circ C) \]

\[ c = 0.904 \, J/g\cdot ^\circ C \]
Honors Heating Curve Race

1. Label the energy on the graph at each section. (Use the space on the diagram if a formula is needed.)
2. Calculate the energy absorbed at each part of the heating curve. Show your work. Round at the end.

Specific heat of \( \text{H}_2\text{O}(s) = 2.03 \text{ J/g°C} \)
Specific heat of \( \text{H}_2\text{O}(l) = 4.184 \text{ J/g°C} \)
Specific heat of \( \text{H}_2\text{O}(g) = 2.01 \text{ J/g°C} \)

\( \Delta H_{\text{fus}} = 6.01 \text{ kJ/mol} \)
\( \Delta H_{\text{vap}} = 40.7 \text{ kJ/mol} \)

**Ice (H\(_2\)O(s) From -50.0°C to 0.0°C)**

\[
q = m \cdot C \Delta T
\]
\[
q = (100.0 \text{ g}) (2.03 \text{ J/g°C}) (50.0°C)
\]
\[
q = 10150 \text{ J}
\]

Convert the answer from Joules to kilojoules (kJ)

\[
\frac{10150 \text{ J}}{1000 \text{ J}} = 10.150 \text{ kJ}
\]

**Melting (0.0°C)**

\[
\Delta H_{\text{fus}} = 6.01 \text{ kJ/mol}
\]
\[
q = \frac{m \cdot C \Delta T}{100.0 \text{ g} \cdot 1 \text{ mol}} = \frac{6.01 \text{ kJ}}{18.02 \text{ g} \cdot 1 \text{ mol}}
\]
\[
q = 33.352 \text{ kJ}
\]

**Water (H\(_2\)O(l) From 0.0°C to 100.0°C)**

\[
q = m \cdot C \Delta T
\]
\[
q = 100 \text{ g} (4.184 \text{ J/g°C}) (100.0°C)
\]
\[
q = 41840 \text{ J}
\]

Convert the answer from Joules to kilojoules (kJ)

\[
\frac{41840 \text{ J}}{1000 \text{ J}} = 41.840 \text{ kJ}
\]

**Vaporization (100.0°C)**

\[
\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}
\]
\[
q = \frac{18.02 \text{ g} \cdot 1 \text{ mol}}{100.0 \text{ g} \cdot 1 \text{ mol}} \cdot 40.7 \text{ kJ}
\]
\[
q = 225.860 \text{ kJ}
\]

**Steam (H\(_2\)O(g) From 100.0°C to 110.0°C)**

\[
q = m \cdot C \Delta T
\]
\[
q = \frac{100 \text{ g} \cdot (0°C - 100°C)}{1 \text{ mol}} (2.01 \text{ J/g°C})
\]
\[
q = 2010.0 \text{ J}
\]

Convert the answer from Joules to kilojoules (kJ)

\[
\frac{2010.0 \text{ J}}{1000 \text{ J}} = 2.010 \text{ kJ}
\]

Add the energies in kJ from each box together.

\[
10.2 \text{ kJ} + 225.9 \text{ kJ} + 33.35 \text{ kJ} + 2.010 \text{ kJ} + 41.84 \text{ kJ} = 313.22 \text{ kJ}
\]

W/sf \[\boxed{313.2 \text{ kJ}}\]
Phase Changes for Water

The graph shows the relationship between temperature and heat energy during the phase changes of water. Study the graph and answer the questions.

1. Why doesn't the temperature increase during melting?
   The heat is "hidden" latent energy given in to change the crystal to liquid by moving particles.

2. Is energy required for each phase change?
   Yes

3. Can both liquid water and steam exist at 100°C?
   Yes

4. What must be changed, temperature or heat energy, during condensation?
   Heat

5. How would you describe the change in the arrangement of particles as heat energy and temperature increase?
   More random than arranged

6. What rule can you state about the relationship between phase changes and temperature?
   Phase change - No temp change
   Temp change - No phase change

7. What rule can you state about the relationship between phase changes and heat energy?
   For heat energy → phase change
Energy Transformations:

Heat always flows from **hot** objects to **cold** objects.

A flask where an endothermic reaction is occurring feels **cold** whereas a flask where an exothermic reaction is occurring feels **warm**.

**Calculating Heat**  \( Q = m(\Delta T) C_p \)

Ex: Calculate the amount of energy required to completely boil 100.0 mL of water that you get from the faucet at a temperature of 22.5 °C.

\[
\Delta H = (100.0 \text{ g})(100^\circ \text{C} - 22.5^\circ \text{C})(4.184 \text{ J/g}^\circ \text{C}) \\
Q = \text{32400 J}
\]

**Calorimetry:**

- \( q_{\text{sys}} = q_{\text{surr}} \)
- \( m(\Delta T) C_p = m(\Delta T) C_p \)

**Changes of State:** (Is each endo or exo?? Give the correct sign for \( \Delta H \))

- Heat of Fusion, \( \Delta H = (+) \)
- Heat of Solidification, \( \Delta H = (-) \)
- Heat of Vaporization, \( \Delta H = (+) \)
- Heat of Condensation, \( \Delta H = (-) \)

Give the sign for \( \Delta H \) and tell whether the reaction is endo or exothermic:

\[
\begin{align*}
\text{H}_2\text{O}(l) &\rightarrow \text{H}_2\text{O}(g) \quad \Delta H = + \quad \text{endo} \\
\text{H}_2\text{O}(l) &\rightarrow \text{H}_2\text{O}(s) \quad \Delta H = - \quad \text{exo}
\end{align*}
\]

**Thermochemical Equations:**

Standard conditions for thermo eqns: **1 atm** and **25°C (298K)**

- In any reaction, heat is either released or absorbed – heat can be written as reactant or product

Ex: \( \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} \quad \Delta H = -154 \text{ KJ} \)

Ex: \( 2\text{Fe}(s) + 3/2\text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s) + 824.2 \text{ kJ} \)

**Hess’ Law**

**Standard Enthalpy (Heat) of Formation**

Calculate the \( \Delta H_{\text{rxn}} \) using standard heats of formation from Table C-13.

**Vocabulary:**

Specific heat (with units), Heat capacity (with units), Calorimeter, Law of conservation of energy

**Problems:**

pp. 524-525
46, 48, 49, 55, 60, 62, 75, 76, 79, 80, 82, 83, 84, 85, 88, 89 a&b
46) Temperature is a measurement of heat. Heat is energy that is in the process of flowing from a warmer object to a cooler object.

48) 1 Calorie = 1000 cal

49) specific heat = \( J/\text{g}\degree \text{C} \)

50) \( \Delta H_{\text{exo}} = -, \Delta H_{\text{endo}} = + \)

60) a) \( \Delta H = + \)
    b) \( \Delta H = - \)
    c) \( \Delta H = + \)
    d) \( \Delta H = - \)

62) \( \sum \Delta H_{\text{products}} = \sum \Delta H_{\text{reactants}} \)

75) \[ \frac{544 \text{ Calories}}{1000 \text{ cal}} = \frac{544000 \text{ cal}}{1 \text{ Calorie}} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 2280000 \text{ J} \]

76) \[ \frac{138 \text{ KJ}}{1000 \text{ J}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 33000 \text{ cal} = 3.30 \times 10^4 \text{ cal} \]

79) \[ Q = m \Delta T \cdot C_p \]
\[ Q = \frac{2000 \text{ kg}}{1000 \text{ g}} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times (29\degree \text{C} - 10\degree \text{C}) (0.803 \frac{\text{J}}{\text{g}\degree \text{C}}) \]
\[ Q = 30,000,000 \text{ J} \]

80) \[ Q = m \Delta T \cdot C_p \]
\[ Q_m = 63.3 \text{ g} \times (292.4 \text{ K} - 384.4 \text{ K}) (0.24 \frac{\text{J}}{\text{g}\degree \text{C}}) \]
\[ Q_m = -1397.16 \text{ J} \]
\[ Q_w = 1397.16 \text{ J} = m (292.4 \text{ K} - 290.0 \text{ K}) (4.18 \frac{\text{J}}{\text{g}\degree \text{C}}) \]
\[ m = 140 \text{ g} \]
82) \[ \frac{343 \text{ g C}_2 \text{H}_5 \text{OH}}{1 \text{ mol}} = 38.6 \text{ kJ} = 287 \text{ kJ} \]
\[ \frac{146.08 \text{ g}}{1 \text{ mol}} \]

83) \[ \frac{1255 \text{ g H}_2 \text{O}}{1 \text{ mol H}_2 \text{O}} = -40.7 \text{ kJ} = -2830 \text{ kJ} \]
\[ \frac{18.02 \text{ g H}_2 \text{O}}{1 \text{ mol}} \]

84) \[ \frac{206 \text{ g H}_2}{1 \text{ mol H}_2} = -286 \text{ kJ} = 29,200 \text{ kJ released} \]
\[ \frac{2.02 \text{ g H}_2}{1 \text{ mol}} \]

85) \[ 5.66 \text{ kJ NH}_3 \]
\[ \frac{1 \text{ mol NH}_3}{1 \text{ mol}} = 17.04 \text{ g NH}_3 \]
\[ 5.66 \text{ kJ} \]

88) \[ \text{Sn(s)} + 2\text{Cl}_2(g) \rightarrow \text{SnCl}_4(l) \]
\[ \text{Sn(s)} + \text{Cl}_2(g) \rightarrow \text{SnCl}_2(s) \quad \Delta H = -325 \text{ kJ} \]
\[ \text{SnCl}_2(s) + \text{Cl}_2(g) \rightarrow \text{SnCl}_4(l) \quad \Delta H = -186 \text{ kJ} \]
\[ \Delta H = -511 \text{ kJ} \]

89 a) \[ 2 \text{NaHCO}_3(s) \rightarrow \text{Na}_2\text{CO}_3(s) + \text{CO}_2(g) + \text{H}_2\text{O}(g) \]
\[ 2(-950.8) \]
\[ -1130.7 \]
\[ -343.509 \]
\[ -241.818 \]

\[ \Delta H = -1766.027 + 1901.6 \]
\[ -135.573 \approx 135.6 \text{ kJ} \]

b) \[ \text{H}_2(g) + 0.25\text{O}_2(g) \rightarrow \text{H}_2\text{O}(l) \]
\[ 0 \quad 0 \quad -187.8 \]

\[ \Delta H = -187.8 \text{ kJ} \]
Honors Study Guide: Ch. 16 – Thermochemistry

Energy Transformations:
Heat *always* flows from warmer objects to cooler objects
A flask where an endothermic reaction is occurring feels cool whereas a flask where an exothermic reaction is occurring feels warm

Calculating Heat  \[ Q = m(\Delta T) C_p \]
Ex: Calculate the amount of energy required to completely boil 100.0 mL of water that you get from the faucet at a temperature of 22.5 °C.
\[ Q = (100.0 \, \text{g})(100.0^\circ \text{C} - 22.5^\circ \text{C})(4.184 \, \text{J/g}^\circ \text{C}) \]
\[ Q = 32,400 \, \text{J} \]

Calorimetry:
- \[ q_{sys} = q_{surr} \]
- \[ m(\Delta T) C_p = m(\Delta T) C_p \]

Changes of State:  (Is each endo or exo?? Give the correct sign for \( \Delta H \))
- Heat of Fusion, \( \Delta H = (+) \)
- Heat of Solidification, \( \Delta H = (-) \)
- Heat of Vaporization, \( \Delta H = (+) \)
- Heat of Condensation, \( \Delta H = (-) \)

Give the sign for \( \Delta H \) and tell whether the reaction is endo or exothermic:
\[ H_2O(l) \rightarrow H_2O(g) \quad \Delta H = + \quad \text{endothermic} \]
\[ H_2O(l) \rightarrow H_2O(s) \quad \Delta H = - \quad \text{exothermic} \]

Thermochemical Equations:
Standard conditions for thermo eqns: 1 atm and 25°C (298K)
- in any reaction, heat is either released or absorbed ~ heat can be written as reactant or product

Ex: \[ \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} \quad \Delta H = -154 \, \text{KJ} \]
Ex: \[ 2\text{Fe}(s) + 3/2\text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s) \quad + \quad 824.2 \, \text{kJ} \]

Hess’ Law

Standard Enthalpy (Heat) of Formation
Calculate the \( \Delta H_{rxn} \) using standard heats of formation from Table C-13

Vocabulary:
Specific heat (with units), Heat capacity (with units), Calorimeter, Law of conservation of energy

Problems: pp. 524-525
46. Temperature is a measurement of heat. Heat is energy that is in the process of flowing from a warmer object to a cooler object.
48. 1 Calorie = 1000 cal
49. specific heat = J/g°C
55. \( \Delta H \) exothermic = −, \( \Delta H \) endothermic = +
60. a. \( \Delta H = + \), b. \( \Delta H = - \), c. \( \Delta H = + \), d. \( \Delta H = - \)
62. \( \Sigma \Delta H_{products} - \Sigma \Delta H_{reactants} \)
75. 544,000 cal, 2,280,000 J
76. 33,000 = 3.30 \times 10^4 \text{cal}
79. \( Q = 30,000,000 \, \text{J} \)
80. \( m = 140 \, \text{g} \)
82. 287 kJ
83. 2830 kJ
84. -29,200 kJ or 29,200 kJ released
85. 17.04 g \[ \text{Fe}_2\text{O}_3(s) \]
88. \( \Delta H = -511 \, \text{kJ} \)
89. a \approx 135.6 \, \text{kJ}; b = -187.8 \, \text{kJ}

\[ C = -176.0 \, \text{kJ}; D = 1080.0 \, \text{kJ} \]
\[ E = -2435.7 \, \text{kJ} \]