

**Learning objective 5.4** The student is able to use conservation of energy to relate the magnitudes of the energy changes occurring in two or more interacting systems, including identification of the systems, the type (heat versus work), or the direction of energy flow. [See SP 1.4, 2.2, connects to Essential knowledge 5.B.1, 5.B.2]

**Learning objective 5.5** The student is able to use conservation of energy to relate the magnitudes of the energy changes when two nonreacting substances are mixed or brought into contact with one another. [See SP 2.2, connects to Essential knowledge 5.B.1, 5.B.2]

**Calorimetry Calculations Learning objective 5.6** The student is able to use calculations or estimations to relate energy changes associated with heating/cooling a substance to the heat capacity, relate energy changes associated with a phase transition to the enthalpy of fusion/vaporization, relate energy changes associated with a chemical reaction to the enthalpy of the reaction, and relate energy changes to  $P\Delta V$  work. [See SP 2.2, 2.3; Essential knowledge 5.B.3]

**LAB: Calorimetry Learning objective 5.7** The student is able to design and/or interpret the results of an experiment in which calorimetry is used to determine the change in enthalpy of a chemical process (heating/cooling, phase transition, or chemical reaction) at constant pressure. [See SP 4.2, 5.1; Essential knowledge 5.B.4] **Enthalpy and Bond Energies**

**Learning objective 5.8** The student is able to draw qualitative and quantitative connections between the reaction enthalpy and the energies involved in the breaking and formation of chemical bonds. [See SP 2.3, 7.1, 7.2; Essential knowledge 5.C.2]

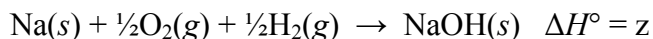
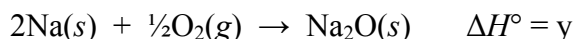
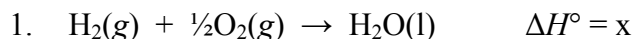
**Learning objective 3.11** The student is able to interpret observations regarding macroscopic energy changes associated with a reaction or process to generate a relevant symbolic and/or graphical representation of the energy changes. [See SP 1.5, 4.4; Essential knowledge 3.C.2]

Equation	Variables	Explanation
$W = P\Delta V$	$W = \text{Work}$ $P = \text{Pressure}$ $\Delta V = \text{Change in Volume}$	This equation is used with gasses. When there is pressure/work/volume involved for gasses, use this equation.
$\Delta E = q + W$	$\Delta E = \text{Change in Energy}$ $q = \text{heat}$ $W = \text{Work (}P\Delta V\text{)}$	First Law of Thermodynamics
$Q = mc\Delta T$	$Q = \text{heat produced or absorbed}$ $m = \text{mass of object or solvent}$ $c = \text{specific heat of object or solvent}$ $\Delta T = \text{change in temperature}$	In problems when you have a reaction happening in a calorimeter, we have $q_{rxn} + q_{sol} = 0 \Rightarrow q_{rxn} = -q_{sol}$ so we can use this expression of $Q$ to $q_{sol}$ to find the heat released or absorbed by the reaction.  In problems when you are putting a metal ball into a liquid and they have different initial temperatures, we know the final temperatures will be the same, and that the amount of heat leaving the hotter object will be equal in magnitude to the heat entering the cooler object, so there is an equivalence you can make.
$Q_{cal} = C\Delta T$	$Q_{cal} = \text{heat change in calorimeter}$ $C = \text{specific heat of calorimeter}$ $\Delta T = \text{change in temperature}$	This applies to a bomb calorimeter question since we know that $q_{rxn} + q_{sol} + q_{bomb} = 0$ $\Rightarrow q_{rxn} = -q_{sol} - q_{bomb}$ $= -mc\Delta T - C\Delta T.$

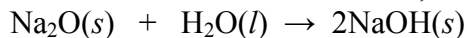
Level 1: #5, 15, 20, 22, 27, 31, 35, 40, 49, 51, 55, 67, 73, 95 from textbook

Level 2 and 3

### Thermochemical Equations and Enthalpy



Based on the information above, what is the standard enthalpy change for the following reaction?



(A)  $x + y + z$

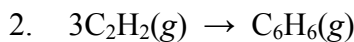
(B)  $x + y - z$

(C)  $x + y - 2z$

(D)  $2z - x - y$

(E)  $z - x - y$

Ans: D



What is the standard enthalpy change,  $\Delta H^\circ$ , for the reaction represented above?

$\Delta H_f^\circ \text{C}_2\text{H}_2(\text{g}) = 230 \text{ kJ/mol}$

$\Delta H_f^\circ \text{C}_6\text{H}_6(\text{g}) = 83 \text{ kJ/mol}$

(A)  $-607 \text{ kJ}$

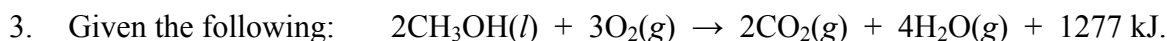
(B)  $-147 \text{ kJ}$

(C)  $-19 \text{ kJ}$

(D)  $+19 \text{ kJ}$

(E)  $+773 \text{ kJ}$

Ans: A



$\Delta H_f^\circ \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$ ;  $\Delta H_f^\circ \text{H}_2\text{O}(\text{g}) = -241.8 \text{ kJ/mol}$ .

a. Calculate the enthalpy of formation of  $\text{CH}_3\text{OH}(\text{l})$  Ans:  $-238.6 \text{ kJ/mol}$

b. How much heat will be evolved when 250.0 mL of methanol is consumed? Assume that the density of liquid methanol is 0.800 g/mL. Ans: 3,985 kJ

c. How many grams of methanol will be required to produce 4750.0 kJ of heat? Ans: 238.4 g

4. The reaction of quicklime (CaO) with water produces slaked lime [ $\text{Ca}(\text{OH})_2$ ], a substance widely used in the construction industry to make mortar and plaster. The reaction of quicklime and water is highly exothermic.



(A) What is the heat of reaction per gram of calcium oxide (CaO) reacted?

**Answer:  $-6.24 \text{ kJ/g CaO}$**

(B) How much heat in kilojoules, is associated with the production of 50.0 kg of slaked lime?

**Answer:  $-2.36 \times 10^5 \text{ kJ}$**

(C) If 10.0 g CaO is added to 100.0 g  $\text{H}_2\text{O}$  at 25 °C in an insulated container, will the temperature of the resulting mixture reach the boiling point of water? [Use a value of 1.09 J/g-°C for the specific heat of  $\text{Ca}(\text{OH})_2(\text{s})$ .]

**Answer: Yes**



Heat of rxn =  $-890.32 \text{ kJ}$ . How many kilojoules will be given off by the burning of 451 g of methane?

Ans:  $2.51 \times 10^4 \text{ kJ}$

6. Given the thermochemical equation  $2\text{Hg}(l) + \text{O}_2(g) \rightarrow 2\text{HgO}(s)$   $\Delta H = -181.66 \text{ kJ}$
- Write the thermochemical equation for the decomposition of 2 mol  $\text{HgO}(s)$  to  $\text{Hg}(l)$  and  $\text{O}_2(g)$ .
  - Write the thermochemical equation that shows the formation of 1 mol  $\text{HgO}$  from  $\text{Hg}(l)$  and  $\text{O}_2(g)$ .

Answer: a)  $2\text{HgO}(s) \rightarrow 2\text{Hg}(l) + \text{O}_2(g)$   $\Delta H = +181.66 \text{ kJ}$   
 b)  $\text{Hg}(l) + 1/2\text{O}_2(g) \rightarrow \text{HgO}(s)$   $\Delta H = -90.83 \text{ kJ}$

7. A certain gas is compressed in a cylinder by a constant pressure of 9.0 atm from a volume of 35.0 L to 4.00 L. Calculate the work ( $w$ ) for this process. Conversion factor =  $8.314 \text{ J}/0.0821 \text{ L}\cdot\text{atm}$

Ans: 28.3 kJ

8.  $\text{H}_2(g) + 1/2\text{O}_2(g) \rightarrow \text{H}_2\text{O}(g)$   $\Delta H = -241.8 \text{ kJ}$   
 (2)  $\text{H}_2(g) + 1/2\text{O}_2(g) \rightarrow \text{H}_2\text{O}(l)$   $\Delta H = -285.8 \text{ kJ}$

On the basis of the above data, which of the following statements is **false**?

- Reaction (1) is exothermic.
- Reaction (2) is the formation reaction for  $\text{H}_2\text{O}(l)$ .
- The reverse of reaction (2) is endothermic.
- The energy content of  $\text{H}_2\text{O}(g)$  is lower than  $\text{H}_2\text{O}(l)$ .
- $\Delta H$  for the reaction:  $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g)$  is  $+44 \text{ kJ/mol}$ .

Ans: D

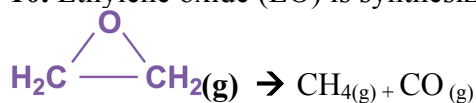
9. Given:  $\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g)$   $\Delta H^\circ_{\text{rxn}} = -393 \text{ kJ}$

How many grams of  $\text{C}(s)$  must burn in this way to release 275 kJ of heat?

Ans; 8.40g

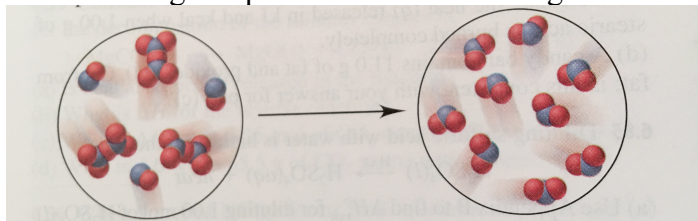
### Calculating $\Delta H^\circ_{\text{rxn}}$ from $\Delta H^\circ_f$

10. Ethylene oxide (EO) is synthesized from ethylene and decomposes explosively upon heating.



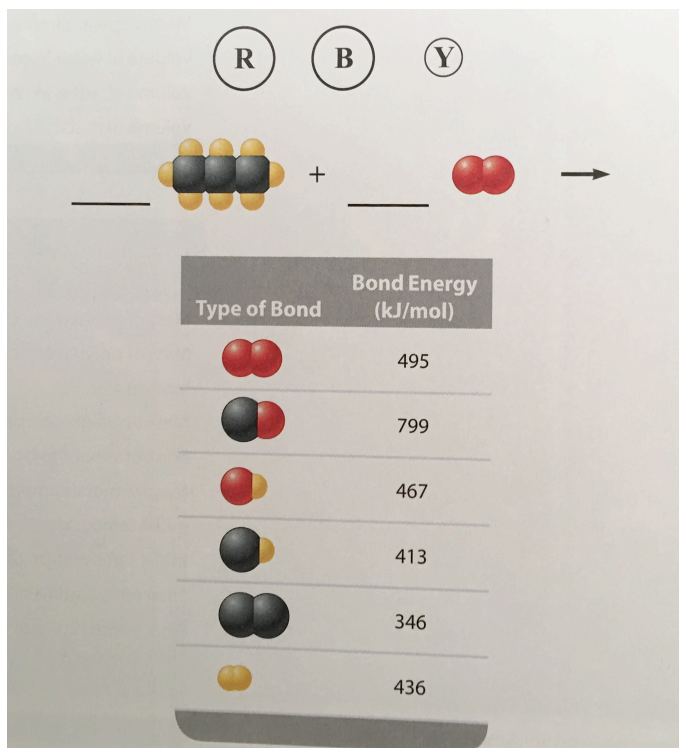
Calculate  $\Delta H^\circ_{\text{rxn}}$  using  $\Delta H^\circ_f$  values from the internet.

11. Following is a particulate level drawing of reaction between N atoms (blue) and O atoms (red) at 298 K.



- Write a balanced equation for the reaction.
- Calculate  $\Delta H^\circ_{\text{rxn}}$  using heats of formation.
- If each molecule of the product represents  $1.50 \times 10^{-2}$  moles, what quantity of heat is absorbed or released?

12. Look at the following particulate level drawing carefully:

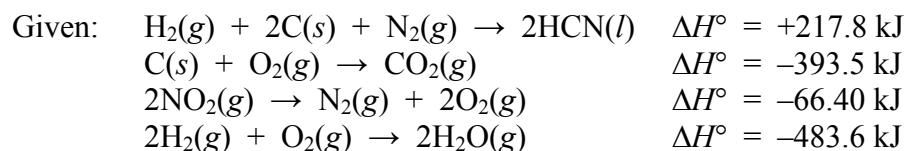
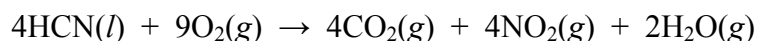


1. Write the balanced chemical reaction represented by the drawing.

Hint: Black atoms C, yellow H, red O  
 2. Calculate  $\Delta H^\circ_{\text{rxn}}$  by using bond enthalpy data given.

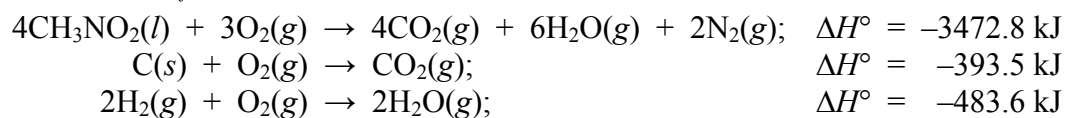
### Hess' Law

13. Use Hess' Law to calculate  $\Delta H^\circ$  for the following reaction:



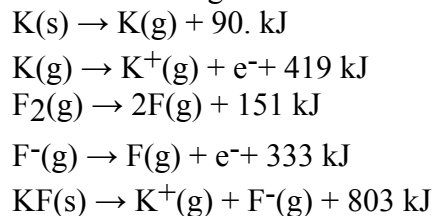
Ans:  $\text{DH}_f$  -2360 kJ

14. Calculate  $\Delta H_f^\circ$  of nitromethane,  $\text{CH}_3\text{NO}_2$ , given the following information:



Ans:  $\text{DH}_f = 112 \text{ kJ/mol}$

12. Given the following data:



Calculate  $\Delta H$  for the reaction  $\text{K}(s) + 1/2\text{F}_2(g) \rightarrow \text{KF}(s)$

Ans: 552 kJ

### Calorimetry

13. 0.500g of magnesium chips are placed in a coffee-cup calorimeter and 100.0 ml of 1.00 M HCl is added to it. The reaction that occurs is:  

$$\text{Mg (s)} + 2\text{HCl (aq)} \rightarrow \text{H}_2 \text{ (g)} + \text{MgCl}_2 \text{ (aq)}$$
 The temperature of the solution increases from 22.2°C (295.4 K) to 44.8°C (318.0 K). What's the enthalpy change for the reaction, per mole of Mg? (Assume specific heat capacity of solution is 4.20 J/(g \* K) and the density of the HCl solution is 1.00 g/ml.)  
 Ans.  $-4.64 \times 10^5 \text{ J/mol Mg}$
14. 200. ml of 0.400 M HCl is mixed with the same amount and molarity of NaOH solution, inside a coffee-cup calorimeter. The temperature of the solutions before mixing was 25.10°C, and 27.78°C after mixing and letting the reaction occur. Find the molar enthalpy of the neutralization of the acid, assuming the densities of all solutions are 1.00 g/ml and their specific heat capacities are 4.20 J/(g \* K).
15. You are working for a fancy laboratory, and your boss has asked to you perform a calorimetry investigation to determine the heat of neutralization when hydrochloric acid and sodium hydroxide are mixed.
- a. Write the complete, balanced molecular equation for the reaction between hydrochloric acid and sodium hydroxide.  
 Ans:  $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
- b. Write the balanced net ionic equation for this reaction.  
 Ans:  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
- c. You measure 50.00mL of each 0.10M solution, put them in separate containers, and let them sit for several minutes. You record the initial temperature for each solution and then mix them in a coffee cup calorimeter. You record the temperature during the entire process. Assuming that the acid and base solutions have the same density and specific heat as water, explain HOW you would calculate the heat produced during the procedure. (You CANNOT actually calculate anything. Simply explain your plan of attack.)  
 Ans: -take initial temp of each solution  
 - take final temp and mass  

$$Q_{\text{rxn}} + q_{\text{solution}} = 0$$

$$Q_{\text{rxn}} + (m_{\text{solution}} \times c_{\text{solution}} \times (t_{\text{final}} - t_{\text{initial}})) = 0$$
- d. Congratulations! The computer is ready to give you the data for your experiment. The solutions started at 23.5°C and the highest recorded temperature was 29.6°C. The density of water is 1.00g/mL and the specific heat of water is 4.20 J/g K. Assuming all of the energy from the reaction went into the solution, how much heat was produced during the procedure?  
 Ans:  $Q = mc\Delta T$   

$$= (100\text{g} \times 4.20\text{g/gC} \times 6.1 \text{ C}) = 2562 \text{ J}$$
- e. Use the net ionic equation from #2 to calculate the number of moles of product formed during this reaction.  
 Ans: Moles = 0.05 x 0.10 M  

$$= 0.005 \text{ moles H}^+ = 0.005 \text{ moles product}$$
- f. What is the heat of neutralization ( $\Delta H$ ) per mole for this reaction?  
 Ans: 0.005 moles  $\rightarrow$  -2,562 KJ  
 1 mol  $\rightarrow$  -512.4 KJ/mol

Your boss is impressed by your work, but would like to you to complete a follow up experiment. She wants you to find the heat of neutralization of nitric acid and potassium hydroxide using the exact same amounts and same concentrations as you did before.

g. Write the complete, balanced molecular equation for the reaction between nitric acid and potassium hydroxide.

Ans:  $\text{KOH} + \text{HNO}_3 \rightarrow \text{KNO}_3 + \text{H}_2\text{O}$

h. Write the net ionic equation for this reaction.

Ans:  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$

i. Given you are using the exact same amounts and concentrations as you did in your first experiment, predict how the calculated value of heat of reaction will compare to the answer you found in #8. Explain your prediction.

Ans: Same; same temp change by same # of moles  $\rightarrow$  same heat of neutralization

j. Oh no! Your experiment was going so well, but then you realized you had set up your calorimeter incorrectly and a significant amount of heat was lost into the air during this experiment. What effect does this error have on your calculated value of  $\Delta H$ ? Justify your answer.

Ans:  $\Delta H$  lower because you lost part of heat to surroundings.

10. A 1.00 M aqueous solution of NaOH, a 0.50 M aqueous solution of  $\text{H}_2\text{SO}_4$ , and a coffee-cup calorimeter were allowed to stand at a room temperature of 25.4 C until the temperature of all three reached 25.4 C. A 50.0-mL sample of the 1.00 M NaOH was then placed in the calorimeter. 50.0 mL of 0.50 M  $\text{H}_2\text{SO}_4$  was added as rapidly as possible, and the two solutions were mixed thoroughly. The temperature rose to 31.9 C. From this experiment, what is the heat of neutralization of one mole of sulfuric acid? For simplicity, assume that the densities of the NaOH and  $\text{H}_2\text{SO}_4$  solutions were 1.00 g/mL and that the specific heat of the solution after reaction was 4.18 J/gC.

Answer: -112 kJ/mol

### Specific Heat

11. A mass of 130. g of a metal is heated to 135 °C and placed in 250. mL of water at 21.3°C. If the final temperature of the water is 45.8 °C, calculate the specific heat of the metal in J/(°C-g).

Ans: 2.21 J/K-g

12. 84.12 g of gold at 120.1 °C is placed in 106.4 g of water at 21.4 °C. At what temperature will the gold and the water reach thermal equilibrium?

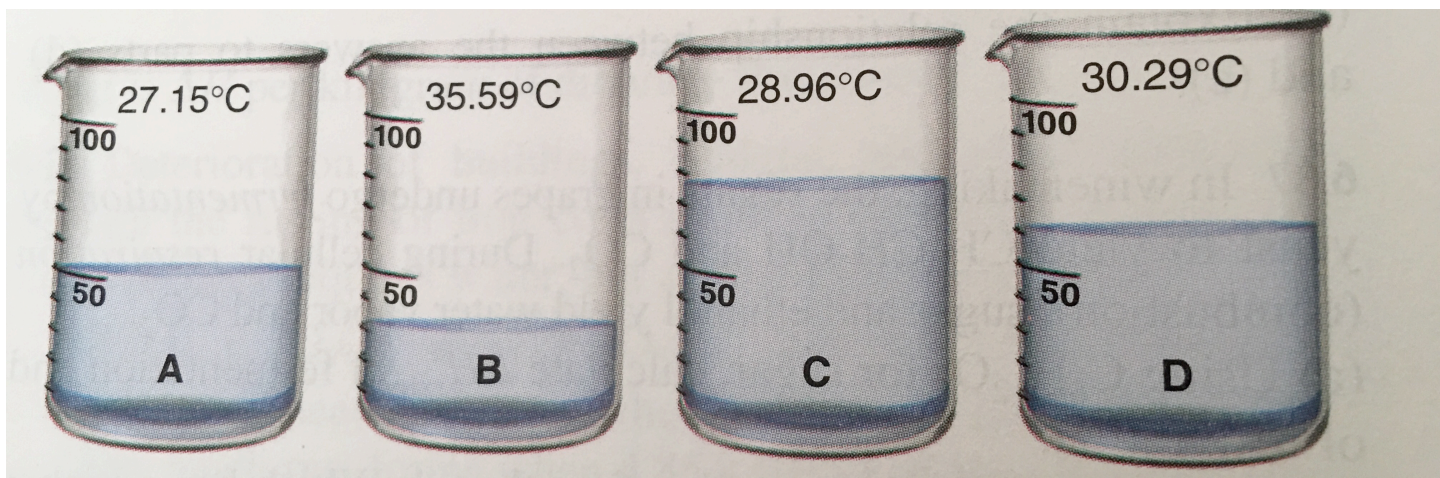
$C_{\text{Au}} = 0.129 \text{ J}/(\text{°C-g})$ ;  $C_{\text{H}_2\text{O}} = 4.184 \text{ J}/(\text{°C-g})$

Ans: 23.8

13. How many joules must be added to a 63.43-g sample of Fe(s) to raise the sample's temperature from 19.7 to 54.2 degrees C? Specific heat of iron is 0.45 J/g. °C.

Ans:  $9.8 \times 10^2 \text{ J}$

14. Four 50.-g samples of different colorless liquids were placed in beakers with initial temperature of 25°C. Each liquid is heated until 450. J of heat has been absorbed. Rank the liquids in order of their specific heat capacity.



Credit Zumdahl

15. The specific heat of liquid water is  $4.18 \text{ J/}^\circ\text{C}\cdot\text{g}$  and the specific heat of carbon is  $0.71 \text{ J/}^\circ\text{C}\cdot\text{g}$ . A  $10.0 \text{ g}$  sample of water and a  $10.0 \text{ g}$  sample of Carbon are each subjected to  $155 \text{ J}$  of heat. If both samples started at  $25^\circ\text{C}$ , which substance will have the higher final temperature, and by what magnitude?

A) Neither. They will have the same final temperature because they started at the same temperature and were exposed to the same quantity of heat.

B) Water, by  $28.7^\circ\text{C}$

C) Carbon, by  $18.1^\circ\text{C}$

D) Water, by  $3.47^\circ\text{C}$

E) Carbon, by  $215^\circ\text{C}$

Ans: C

