

loss to calorimeter container and surroundings, but a correction factor should be used to take into account pressure changes

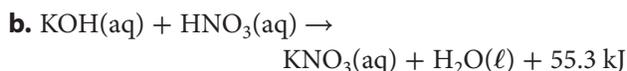
- A and B: heat transfer to water calculated using $Q = mc\Delta T$
- A and C: both use a stirrer to keep water at uniform temperature
- B and C: both are suitable for determining enthalpy of combustion reactions
- A, B, and C: calorimetry is based on the 1st and 2nd laws of thermodynamics; calorimeters are used to measure release or absorption of thermal energy, usually for the purpose of determining the enthalpy change associated with chemical and physical processes

- 8. a.** Equimolar amounts are used and the reaction requires 1:1 ratios of reactants. Therefore the amount of either reactant can be used to determine the molar enthalpy of neutralization.

$$\begin{aligned} Q &= mc\Delta T \\ &= 200.0 \text{ g} \times 4.19 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \times 0.33^\circ\text{C} \\ &= 276.54 \text{ J} \\ &= 0.27654 \text{ kJ} \end{aligned}$$

$$\begin{aligned} n &= cV \\ &= 0.0500 \frac{\text{mol}}{\text{L}} \times 100.00 \text{ mL} \frac{1 \text{ L}}{1000 \text{ mL}} \\ &= 0.00500 \text{ mol} \end{aligned}$$

$$\begin{aligned} \Delta H_{\text{neut}} &= \frac{-0.27654 \text{ kJ}}{0.00500 \text{ mol}} \\ &= -55.3 \frac{\text{kJ}}{\text{mol}} \end{aligned}$$



- 9.** The first and second laws of thermodynamics apply to calorimetry as follows: The first law of thermodynamics states that the energy of the universe is constant. This law applies to calorimetry calculations because we assume that the heat change to the system is equal and opposite compared with the heat change to the surroundings. (The first law holds true even if there is heat lost outside the immediate surrounding but the more heat transferred beyond the calorimeter, the greater the experimental error.) The second law of thermodynamics states that thermal energy spontaneously flows from an object at a higher temperature to one at a lower temperature, when they are in thermal contact. When objects at different temperatures are in contact, thermal energy is transferred from the warmer object to the cooler object

until they are both at the same temperature. The energy change from a chemical reaction, therefore, results in a predictable temperature change to the surroundings, and insulating material can be used to create small, measurable surroundings.

Student diagrams could show a simple calorimeter including the following labels: system; surroundings; effective surroundings, insulating material

- 10.** A flame calorimeter is an open system because both matter and energy can be exchanged with the surroundings.

- 11.** Thermal energy gained by the calorimeter

$$\begin{aligned} Q &= C\Delta T \\ &= 11.6 \frac{\text{kJ}}{^\circ\text{C}} \times 8.41^\circ\text{C} \\ &= 97.556 \text{ J} \end{aligned}$$

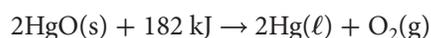
Thermal energy released by burning = -97.556 kJ

$$Q = \frac{-97.556 \text{ kJ}}{6.60 \text{ g}} = -14.7812 \text{ kJ/g}$$

The thermal energy released per gram of biscuit is 14.78 kJ/g

- 12.** Student reports should describe the design of the room, how it is insulated from outside energy changes, how changes in air temperature can be monitored, how $\text{O}_2(\text{g})$ uptake and $\text{CO}_2(\text{g})$ production are measured, how long the subject would stay in the room, what kind of food is used, what exercise is expected by the subject, and how is the mass of the person and/or changes in mass will be monitored. Reference(s) must be included.

- 13. Sample answer:**



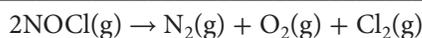
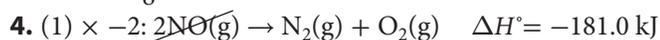
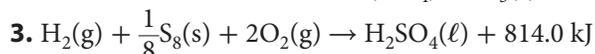
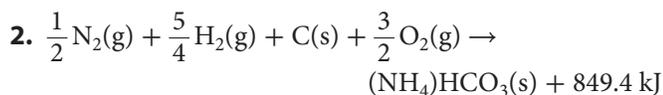
When 182 kJ of thermal energy is added to 2 mol of solid mercury(II) oxide, 2 mol of liquid mercury and 1 mol of oxygen gas are produced. Alternatively, you could say: To decompose 2 mol of mercury(II) oxide, 182 kJ of thermal energy must be added. 2 mol of liquid mercury and 1 mol of oxygen gas are produced.

Answers to Section 5.3 Review Questions (Student textbook page 324)

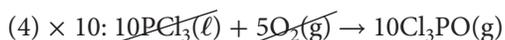
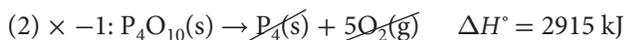
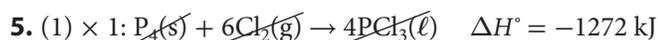
- 1.** The standard enthalpy change for a reaction in which a certain compound is a product is the enthalpy change for that reaction as written. It is applicable for production of that compound in the amount that corresponds to the stoichiometric coefficient in the chemical equation. The equation is:

$$\Delta H_{\text{r}}^\circ = [\Sigma(n\Delta H_{\text{f}}^\circ \text{products})] - [\Sigma(n\Delta H_{\text{f}}^\circ \text{reactants})]$$

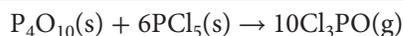
The standard molar enthalpy of formation of a compound represents the enthalpy change when 1 mol of a compound is formed from its elements under standard conditions.



$$\Delta H_r = (-181.0 \text{ kJ}) + (78.0 \text{ kJ}) = -103.0 \text{ kJ}$$



$$\Delta H^\circ = -2670 \text{ kJ}$$



$$\Delta H_r = (-1272 \text{ kJ}) + (2915 \text{ kJ}) + (750 \text{ kJ}) + (-2670 \text{ kJ}) = -277 \text{ kJ}$$

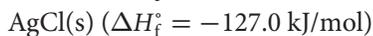
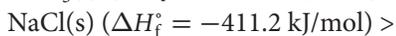
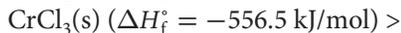
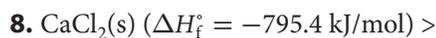
$$6. n_{\text{Ba}} = \frac{5.48 \text{ g}}{137.33 \text{ g/mol}} = 0.03990 \text{ mol}$$

$$\frac{-21.9 \text{ kJ}}{0.03990 \text{ mol}} = \frac{x}{1 \text{ mol}}$$

$$x = -548.87 \text{ kJ}$$

Ba(s) and BaO(s) are equimolar; therefore, the enthalpy of formation for BaO(s) is -549 kJ/mol

7. $\text{O}_2(\text{g})$ is the standard state of this element and is assigned a zero value. To produce $\text{O}(\text{g})$, energy must be added in order to break the strong covalent bonds in the oxygen molecules:



Decomposition is the reverse reaction of formation.

The more negative the ΔH_f° , the more energy will be required to decompose the compound. The four compounds are ranked from the most negative ΔH_f° to the least negative.

$$9. \text{a. } \Delta H_r^\circ = \Sigma(n\Delta H_f^\circ \text{ products}) - \Sigma(n\Delta H_f^\circ \text{ reactants}) \\ = [(9 \text{ mol})(\Delta H_f^\circ \text{ Fe}(\text{s})) + (4 \text{ mol})(\Delta H_f^\circ \text{ Al}_2\text{O}_3(\text{s}))] - \\ [(3 \text{ mol})(\Delta H_f^\circ \text{ Fe}_3\text{O}_4(\text{s})) + (8 \text{ mol})(\Delta H_f^\circ \text{ Al}(\text{s}))]$$

$$= [(9 \text{ mol})(0 \text{ kJ/mol}) + (4 \text{ mol})(-1675.7 \text{ kJ/mol})] - \\ [(3 \text{ mol})(-1118.4 \text{ kJ/mol}) + (8 \text{ mol})(0 \text{ kJ/mol})] \\ = (-6702.8 \text{ kJ}) + (3355.2 \text{ kJ}) = -3347.6 \text{ kJ}$$

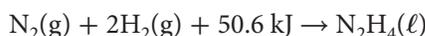
$$\text{b. } n_{\text{Fe}(\text{s})} = \frac{1\,000\,000 \text{ g}}{55.85 \text{ g/mol}} = 1.7905 \times 10^4 \text{ mol} \\ \frac{-3347.6 \text{ kJ}}{9 \text{ mol Fe}(\text{s})} = \frac{x}{1.7905 \times 10^4 \text{ mol Fe}(\text{s})} \\ x = -6.65986 \times 10^6 \text{ kJ}$$

The quantity of heat given off would be $6.660 \times 10^6 \text{ kJ}$

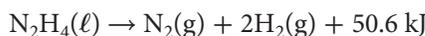
10. *Sample answer:* Hi. Is hydrazine a thermally stable compound? Let's look at the enthalpies involved:

$$\Delta H_f^\circ \text{ N}_2\text{H}_4(\ell) = +50.6 \text{ kJ/mol}$$

The chemical equation for the formation of hydrazine is:



The equation for the thermal decomposition of hydrazine is the opposite:



What this means is that hydrazine releases thermal energy when it decomposes. This means that the bonds in hydrazine are not as strong as the bonds in nitrogen and hydrogen combined. So hydrazine is not very thermally stable.

$$11. \Delta H_r^\circ = \Sigma(n\Delta H_f^\circ \text{ products}) - \Sigma(n\Delta H_f^\circ \text{ reactants})$$

$$\Delta H_r^\circ = [(1 \text{ mol})(\Delta H_f^\circ \text{ N}_2(\text{g})) + (6 \text{ mol})(\Delta H_f^\circ \text{ HF}(\text{g})) \\ + (1 \text{ mol})(\Delta H_f^\circ \text{ Cl}_2(\text{g}))] - [(2 \text{ mol})(\Delta H_f^\circ \text{ ClF}_3(\text{g})) + \\ (2 \text{ mol})(\Delta H_f^\circ \text{ NH}_3(\text{g}))]$$

$$-1200 \text{ kJ} = [(1 \text{ mol})(0 \text{ kJ/mol}) + (6 \text{ mol}) \\ (-273.3 \text{ kJ/mol}) + (1 \text{ mol})(0 \text{ kJ/mol})] - [(2 \text{ mol}) \\ (\Delta H_f^\circ \text{ ClF}_3(\text{g})) + (2 \text{ mol})(-45.9 \text{ kJ/mol})]$$

$$-1200 \text{ kJ} = [(-1639.8 \text{ kJ})] - [(2 \text{ mol})(\Delta H_f^\circ \text{ ClF}_3(\text{g})) \\ + (91.8 \text{ kJ})]$$

$$(2 \text{ mol})(n\Delta H_f^\circ \text{ ClF}_3(\text{g})) = -1731.6 \text{ kJ} + 1200 \text{ kJ}$$

$$(\Delta H_f^\circ \text{ ClF}_3(\text{g})) = -266 \text{ kJ/mol}$$

12. A positive ΔH_f° indicates that the system has gained energy in the formation of the compound. A greater amount of energy is absorbed in breaking bonds in the reactants than is liberated in the formation of bonds in the compound. Student diagrams should resemble Figure 5.13B on page 294 of the student textbook.

$$13. n_{\text{CaCO}_3(\text{s})} = \frac{500.0 \text{ g}}{100.09 \text{ g/mol}} = 4.99550 \text{ mol}$$

$$\Delta H_f^\circ \text{ CaCO}_3(\text{s}) = -1207.6 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ CaO}(\text{s}) = -634.9 \text{ kJ/mol}$$

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