

Chapter 6:

Thermochemistry

Circle one: This reaction is

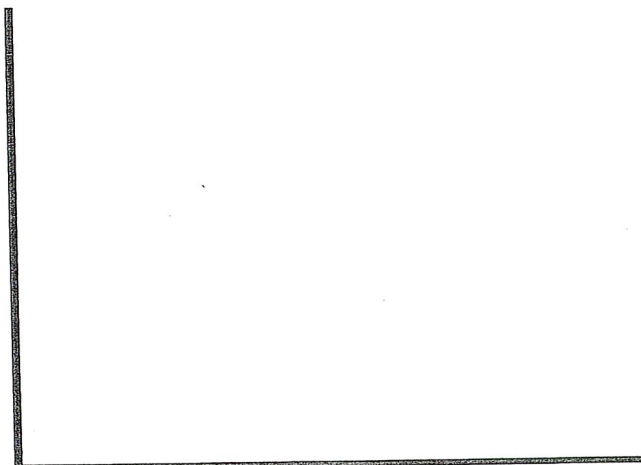
Endothermic

Exothermic

Why?

Sketch a graph of this reaction:

Title: _____



1. Title the graph
2. Label the axis
3. Label the reactants
4. Label the products
5. Label the change in enthalpy with units



Sketch a graph of this reaction:

Title: _____



6. Title the graph
7. Label the axis
8. Label the reactants
9. Label the products
10. Label the change in enthalpy with units

Change in enthalpy (ΔH) is calculated by subtracting the final enthalpy

$$\Delta H = H_P - H_R$$

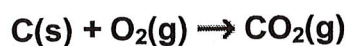
How can you use this formula to support your designation endo/exothermic nature of your graphs?

Name: _____ Class: _____

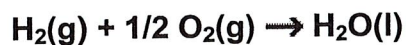
Hess's Law Worksheet

Given the following equations and ΔH° values, determine the heat of the following reactions. Then sketch an enthalpy diagram. (x, y axis, products reactant enthalpy)

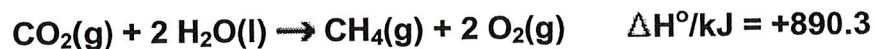
1.



$$\Delta H^\circ/\text{kJ} = -393.5$$



$$\Delta H^\circ/\text{kJ} = -285.8$$



$$\Delta H^\circ/\text{kJ} = +890.3$$



2.



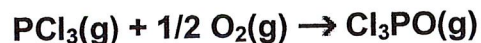
$$\Delta H^\circ/\text{kJ} = -306.4$$



$$\Delta H^\circ/\text{kJ} = -2967.3$$



$$\Delta H^\circ/\text{kJ} = -84.2$$



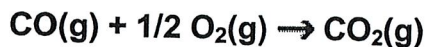
$$\Delta H^\circ/\text{kJ} = -285.7$$



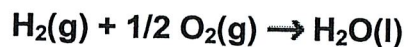
3.



$$\Delta H^\circ/\text{kJ} = -110.5 \text{ kJ}$$



$$\Delta H^\circ/\text{kJ} = -282.9 \text{ kJ}$$



$$\Delta H^\circ/\text{kJ} = -285.8 \text{ kJ}$$



$$\Delta H^\circ/\text{kJ} = -74.8 \text{ kJ}$$



4. Define the following terms:

a) enthalpy

b) exothermic

c) First Law of thermodynamics

e) calorimetry

1. Water has a specific heat capacity of $4.184 \text{ J/g}\cdot^{\circ}\text{C}$.
This means it takes 4.184 J to heat 1.00 gram of water 1.00°C .
 - a) How much energy will it take to heat 10.0 grams of water 1°C ? _____
 - b) How much energy is needed to heat $30.0 \text{ g H}_2\text{O}$ from 10.0°C to 50.0°C ? _____

2. Let's try a standard **calorimetry** problem.
A pot of water (2.5 Liters of water) initially at 25.0°C is heated to boiling (100.0°C).
How much energy (in J) is needed to heat the water? (The density of water is 1 g/mL .)

What would this amount of heat be in kJ? _____

3. What amount of heat is *released* when 175 g of water *cools* from 100.0°C to room temperature, 20.0°C ?
4. We don't always have to warm up or cool down water. The specific heat capacity of copper metal is $0.39 \text{ J/g}\cdot^{\circ}\text{C}$. It is _____ (easier/more difficult) to heat up copper than to heat up water.

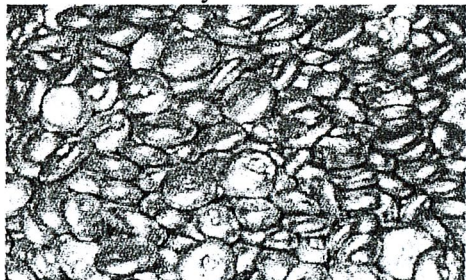
How much energy would it take to heat up a 5.20 g sample of copper from 20.0°C to 100.0°C ?

5. If $300. \text{ J}$ of heat energy were used to heat up a 5.00 gram sample of copper metal and a 5.00 gram sample of water both starting at 10.0°C , calculate the final temperature of each sample?

b. Set up an expression for the energy **absorbed** (q) by the cold water ($\Delta q_{\text{cold}} = m_{\text{cold}} C \Delta T_{\text{cold}}$)

c. Knowing that the heat released = $-$ heat absorbed, combine the two expressions and solve for x .

7. We don't always have to use water. Let's use some **aluminum shot**.



"shot" are these little pellets.

175 grams of hot aluminum ($100.^{\circ}\text{C}$) is dropped into an insulated cup that contains 40.0 mL of ice cold water (0.0°C). Follow the example above to determine the final temperature, x .

a. Set up an expression for the heat lost by the aluminum ($C=0.900 \text{ J/g} \cdot ^{\circ}\text{C}$)

b. Set up an expression for the heat gained by the cold water.

c. Put the two expressions together (don't forget to change one of the signs) and solve for x .

8. Somewhat Confusing Definitions:

There are several terms used in this chapter that sound very similar. Use the data provided to calculate each of them to clarify the differences. I've added some "Notes" that I hope will help.

74.8 J of heat is required to raise the temperature of 18.69 g of silver from 10.0°C to 27.0°C .

a. What is the **heat capacity** of the silver sample? ($\text{J}/^{\circ}\text{C}$)

Note: This is a useful value only for this specific sample of silver.

b. What is the **specific heat capacity** of silver? ($\text{J/g} \cdot ^{\circ}\text{C}$)

Note: This is a useful value for **any** sample of silver that is heated or cooled. This is equivalent to the $4.184 \text{ J} \cdot \text{g}^{-1} \cdot ^{\circ}\text{C}^{-1}$ that we use for water. This value is also called the **specific heat**.

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BL Thermo

Section 5.5 The amount of heat transferred between the system and the surroundings is measured experimentally by **calorimetry**. A **calorimeter** measures the temperature change accompanying a process. The temperature change of a calorimeter depends on its **heat capacity**, the amount of heat required to raise its temperature by 1 K. The heat capacity for 1 mol of a pure substance is called its **molar heat capacity**; for 1 g of the substance, we use the term **specific heat**. Water has a very high specific heat, 4.18 J/g·K. The amount of heat, q , absorbed by a substance is the product of its specific heat, its mass, and its temperature change: $q = (\text{specific heat}) \times (\text{grams of substance}) \times \Delta T$.

If a calorimetry experiment is carried out under a constant pressure, the heat transferred provides a direct measure of the enthalpy change of the reaction. Constant-volume calorimetry is carried out in a vessel of fixed volume called a **bomb calorimeter**. Bomb calorimeters are used to measure the heat evolved in combustion reactions. The heat transferred under constant-volume conditions is equal to ΔE . However, corrections can be applied to ΔE values to yield enthalpies of combustion.

Section 5.6 Because enthalpy is a state function, ΔH depends only on the initial and final states of the system. Thus, the enthalpy change of a process is the same whether the process is carried out in one step or in a series of steps. **Hess's law** states that if a reaction is carried out in a series of steps, ΔH for the reaction will be equal to the sum of the enthalpy changes for the steps. We can therefore calculate ΔH for any process, as long as we can write the process as a series of steps for which ΔH is known.

Section 5.7 The **enthalpy of formation**, ΔH_f° , of a substance is the enthalpy change for the reaction in which the substance is formed from its constituent elements. The **standard enthalpy** of a reaction, ΔH° , is the enthalpy change when all reactants and products are at 1 atm pressure and a specific temperature, usually 298 K (25°C). Combining these ideas, the **standard enthalpy of formation**, ΔH_f° , of a substance is the change in enthalpy for the reaction that forms 1 mol of the substance from its elements with all reactants and products at 1 atm pressure and usually 298 K. For any element in its most stable state at 298 K and 1 atm pressure, $\Delta H_f^\circ = 0$. The standard enthalpy change for any reaction can be readily calculated from the standard enthalpies of formation of the reactants and products in the reaction:

$$\Delta H_{\text{rxn}}^\circ = \sum n \Delta H_f^\circ(\text{products}) - \sum m \Delta H_f^\circ(\text{reactants})$$

Section 5.8 The **fuel value** of a substance is the heat released when 1 g of the substance is combusted. Different types of foods have different fuel values and differing abilities to be stored in the body. The most common fuels are hydrocarbons that are found as **fossil fuels**, such as **natural gas**, **petroleum**, and **coal**. Coal is the most abundant fossil fuel, but the sulfur present in most coals causes air pollution. Coal gasification is one possible way to use existing resources as sources of cleaner energy. Sources of **renewable energy** include solar energy, wind energy, biomass, and hydroelectric energy. These energy sources are essentially inexhaustible and will become more important as fossil fuels are depleted.

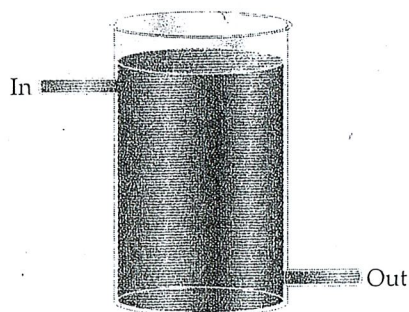
Exercises

The Nature of Energy

- 5.1 In what two ways can an object possess energy? How do these two ways differ from one another?
- 5.2 Suppose you toss a tennis ball upward. (a) Does the kinetic energy of the ball increase or decrease as it moves higher? (b) What happens to the potential energy of the ball as it moves higher? (c) If the same amount of energy were imparted to a ball the same size as a tennis ball, but of twice the mass, how high would it go in comparison to the tennis ball? Explain your answers.
- 5.3 (a) Calculate the kinetic energy in joules of a 45-g golf ball moving at 61 m/s. (b) Convert this energy to calories. (c) What happens to this energy when the ball lands in a sand trap?
- 5.4 (a) What is the kinetic energy in joules of a 950-lb motorcycle moving at 68 mph? (b) By what factor will the kinetic energy change if the speed of the motorcycle is decreased to 34 mph? (c) Where does the kinetic energy of the motorcycle go when the rider brakes to a stop?
- 5.5 In much engineering work it is common to use the British thermal unit (Btu). A Btu is the amount of heat required to raise the temperature of 1 lb of water by 1°F. Calculate the number of joules in a Btu.
- 5.6 A watt is a measure of power (the rate of energy change) equal to 1 J/s. Calculate the number of joules in a kilowatt-hour.
- 5.7 An adult person radiates heat to the surroundings at about the same rate as a 100-watt electric incandescent light bulb. What is the total amount of energy in kcal radiated to the surroundings by an adult in 24 hours?
- 5.8 Describe the source of the energy and the nature of the energy conversions involved when a 100-watt electric lightbulb radiates energy to its surroundings. Compare this with the energy source and energy conversions involved when an adult person radiates energy to the surroundings.
- 5.9 Suppose that a pellet is shot from an air gun straight up into the air. Why does the pellet eventually stop rising and fall back to Earth, rather than simply moving out into

space? In principle, could the pellet ever move out into space?

- 5.10 A bowling ball is dropped from a 100-ft-high tower on Earth. Compare the change in potential energy that it undergoes with dropping the same ball from a 100-ft-high tower on the Moon.
- 5.11 (a) What is meant by the term *system* in thermodynamics? (b) What is special about a closed system?
- 5.12 In a thermodynamic study a scientist focuses on the properties of a solution in a flask that is arranged as



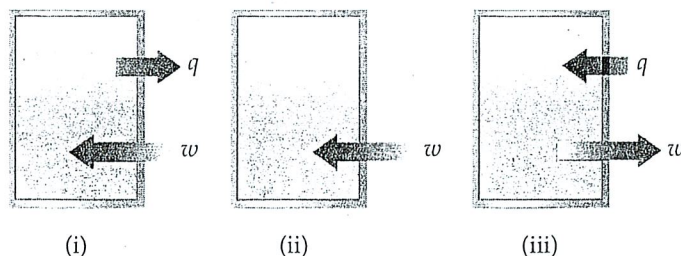
shown in the illustration. A solution is continuously flowing into the flask at the top and out at the bottom such that the amount of solution in the flask is constant in time. (a) Is the solution in the flask a closed system? Why or why not? (b) If it is not a closed system, what could be done with the arrangement in the figure to make a closed system?

- 5.13 (a) What is work? (b) How do we determine the amount of work done, given the force associated with the work?
- 5.14 (a) Not so long ago it was widely believed that heat is not a form of energy. What arguments can you give to convince someone that it is? (b) Under what conditions is heat transferred from one object to another?
- 5.15 Identify the force present, and explain whether work is being performed in the following cases: (a) You lift a pencil off the top of a desk. (b) A spring is compressed to half its normal length.
- 5.16 Identify the force present, and explain whether work is done when (a) a positively charged particle moves in a circle at a fixed distance from a negatively charged particle; (b) an iron nail is pulled off a magnet.

The First Law of Thermodynamics

- 5.17 (a) State the first law of thermodynamics. (b) What is meant by the *internal energy* of a system? (c) By what means can the internal energy of a system increase?
- 5.18 (a) Write an equation that expresses the first law of thermodynamics. (b) In applying the first law, do we need to measure the internal energy of a system? Explain. (c) Under what conditions will the quantities q and w be negative numbers?
- 5.19 Calculate ΔE , and determine whether the process is endothermic or exothermic for the following cases: (a) A system releases 113 kJ of heat to the surroundings and does 39 kJ of work on the surroundings; (b) $q = 1.62$ kJ and $w = -874$ J; (c) the system absorbs 77.5 kJ of heat while doing 63.5 kJ of work on the surroundings.
- 5.20 For the following processes, calculate the change in internal energy of the system, and determine whether the process is endothermic or exothermic: (a) A balloon is heated by adding 900 J of heat. It expands, doing 422 J of work on the atmosphere. (b) A 50-g sample of water is cooled from 30°C to 15°C, thereby losing approximately 3140 J of heat. (c) A chemical reaction releases 8.65 kJ of heat and does no work on the surroundings.
- 5.21 The closed box in each of the following illustrations represents a system, and the arrows show the changes to the system in a process. The lengths of the arrows represent the relative magnitudes of q and w . (a) Which of these processes is endothermic? (b) For which of these processes,

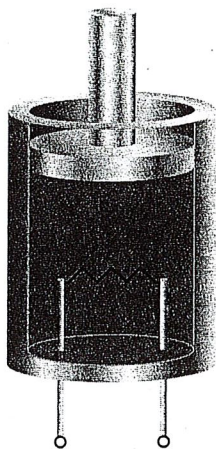
if any, is $\Delta E < 0$? (c) For which process, if any, is there a net gain in internal energy?



- 5.22 A system releases heat to its surroundings and has work done on it by the surroundings. (a) Sketch a box to represent the system, and use arrows to represent the heat and work transferred. (b) Is it possible for ΔE to be positive for this process? Explain. (c) Is it possible for ΔE to be negative for this process? Explain.
- 5.23 A gas is confined to a cylinder fitted with a piston and an electrical heater, as shown in the illustration on the next page. Suppose that current is supplied to the heater so that 100 J of energy are added. Consider two different situations. In case (1) the piston is allowed to move as the energy is added. In case (2) the piston is fixed so that it cannot move. (a) In which case does the gas have the higher temperature after addition of the electrical energy? Explain. (b) What can you say about the values of q and w for each case?

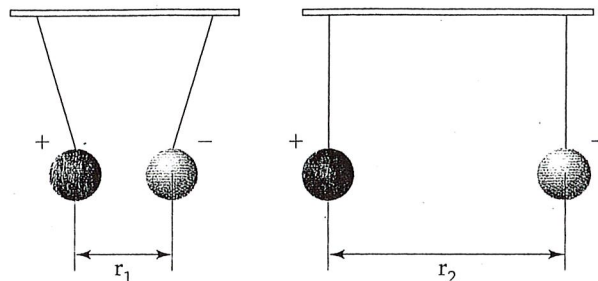
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w in each of these cases? (c) What can you say about the relative values of ΔE for the system (the gas in the cylinder) in the two cases?



- 5.24 Consider a system consisting of two oppositely charged spheres hanging by strings and separated by a distance r_1 , as shown in the illustration in the next column. Suppose they are separated to a larger distance r_2 , by moving them apart along a track. (a) What change, if any, has

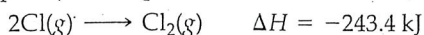
occurred in the potential energy of the system? (b) What effect, if any, does this process have on the value of ΔE ? (c) What can you say about q and w for this process?



- 5.25 (a) What is meant by the term *state function*? (b) Give an example of a quantity that is a state function and one that is not. (c) Is temperature a state function? Why or why not?
- 5.26 Indicate which of the following is independent of the path by which a change occurs: (a) the change in potential energy when a book is transferred from table to shelf; (b) the heat evolved when a cube of sugar is oxidized to $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{g})$; (c) the work accomplished in burning a gallon of gasoline.

Enthalpy

- 5.27 (a) Why is the change in enthalpy a meaningful quantity for many chemical processes? (b) H is a state function, but q is not a state function. Explain. (c) For a given process at constant pressure, ΔH is negative. Is the process endothermic or exothermic?
- 5.28 (a) Under what condition will the enthalpy change of a process equal the amount of heat transferred into or out of the system? (b) Enthalpy is said to be a state function. What is it about state functions that makes them particularly useful? (c) During a constant-pressure process the system absorbs heat from the surroundings. Does the enthalpy of the system increase or decrease during the process?
- 5.29 The complete combustion of acetic acid, $\text{HC}_2\text{H}_3\text{O}_2(\text{l})$, to form $\text{H}_2\text{O}(\text{l})$ and $\text{CO}_2(\text{g})$ at constant pressure releases 871.7 kJ of heat per mole of $\text{HC}_2\text{H}_3\text{O}_2$. (a) Write a balanced thermochemical equation for this reaction. (b) Draw an enthalpy diagram for the reaction.
- 5.30 The decomposition of zinc carbonate, $\text{ZnCO}_3(\text{s})$, into zinc oxide, $\text{ZnO}(\text{s})$, and $\text{CO}_2(\text{g})$ at constant pressure requires the addition of 71.5 kJ of heat per mole of ZnCO_3 . (a) Write a balanced thermochemical equation for the reaction. (b) Draw an enthalpy diagram for the reaction.
- 5.31 Consider the following reaction, which occurs at room temperature and pressure:



Which has the higher enthalpy under these conditions, $2\text{Cl}(\text{g})$ or $\text{Cl}_2(\text{g})$?

- 5.32 Without referring to tables, indicate which of the following has the higher enthalpy in each case: (a) 1 mol $\text{CO}_2(\text{s})$

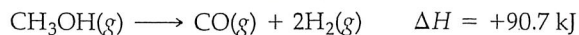
or 1 mol $\text{CO}_2(\text{g})$ at the same temperature; (b) 2 mol of hydrogen atoms or 1 mol of H_2 ; (c) 1 mol $\text{H}_2(\text{g})$ and 0.5 mol $\text{O}_2(\text{g})$ at 25°C or 1 mol $\text{H}_2\text{O}(\text{g})$ at 25°C ; (d) 1 mol $\text{N}_2(\text{g})$ at 100°C or 1 mol $\text{N}_2(\text{g})$ at 300°C .

- 5.33 Consider the following reaction:



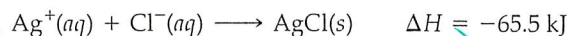
(a) Is this reaction exothermic or endothermic? (b) Calculate the amount of heat transferred when 2.4 g of $\text{Mg}(\text{s})$ reacts at constant pressure. (c) How many grams of MgO are produced during an enthalpy change of -96.0 kJ ? (d) How many kilojoules of heat are absorbed when 7.50 g of $\text{MgO}(\text{s})$ are decomposed into $\text{Mg}(\text{s})$ and $\text{O}_2(\text{g})$ at constant pressure?

- 5.34 Consider the following reaction:



(a) Is heat absorbed or evolved in the course of this reaction? (b) Calculate the amount of heat transferred when 1.60 kg of $\text{CH}_3\text{OH}(\text{g})$ are decomposed by this reaction at constant pressure. (c) For a given sample of CH_3OH , the enthalpy change on reaction is 64.7 kJ. How many grams of hydrogen gas are produced? (d) What is the value of ΔH for the reverse of the previous reaction? How many kilojoules of heat are released when 32.0 g of $\text{CO}(\text{g})$ reacts completely with $\text{H}_2(\text{g})$ to form $\text{CH}_3\text{OH}(\text{g})$ at constant pressure?

- 5.35 When solutions containing silver ions and chloride ions are mixed, silver chloride precipitates:



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- (a) Calculate ΔH for formation of 0.540 mol of AgCl by this reaction. (b) Calculate ΔH for the formation of 1.66 g of AgCl. (c) Calculate ΔH when 0.188 mmol of AgCl dissolves in water.
- 5.36 At one time, a common means of forming small quantities of oxygen gas in the laboratory was to heat KClO_3 :
 $2\text{KClO}_3(s) \longrightarrow 2\text{KCl}(s) + 3\text{O}_2(g) \quad \Delta H = -89.4 \text{ kJ}$
 For this reaction, calculate ΔH for the formation of (a) 4.34 mol of O_2 and (b) 200.8 g of KCl. (c) The decomposition of KClO_3 proceeds spontaneously when it is heated. Do you think that the reverse reaction, the formation of KClO_3 from KCl and O_2 , is likely to be feasible under ordinary conditions? Explain your answer.
- 5.37 You are given ΔH for a process that occurs at constant pressure. What additional information is needed to determine ΔE for the process?
- 5.38 Suppose that the gas phase reaction, $2\text{NO}(g) + \text{O}_2(g) \longrightarrow 2\text{NO}_2(g)$ were carried out in a constant-volume container at constant temperature. Would the measured heat change represent ΔH or ΔE ? If there is a difference, which quantity is larger for this reaction? Explain.
- 5.39 A gas is confined to a cylinder under constant atmospheric pressure, as illustrated in Figure 5.3. When the gas undergoes a particular chemical reaction, it releases 89 kJ of heat to its surroundings and does 36 kJ of P - V work on its surroundings. What are the values of ΔH and ΔE for this process?
- 5.40 A gas is confined to a cylinder under constant atmospheric pressure, as illustrated in Figure 5.3. When 518 J of heat is added to the gas, it expands and does 127 J of work on the surroundings. What are the values of ΔH and ΔE for this process?
- 5.41 Consider the combustion of liquid methanol, $\text{CH}_3\text{OH}(l)$:
 $\text{CH}_3\text{OH}(l) + \frac{3}{2}\text{O}_2(g) \longrightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(l) \quad \Delta H = -726.5 \text{ kJ}$
 (a) What is the enthalpy change for the reverse reaction? (b) Balance the forward reaction with whole-number coefficients. What is ΔH for the reaction represented by this equation? (c) Which is more likely to be thermodynamically favored, the forward reaction or the reverse reaction? (d) If the reaction were written to produce $\text{H}_2\text{O}(g)$ instead of $\text{H}_2\text{O}(l)$, would you expect the magnitude of ΔH to increase, decrease, or stay the same? Explain.
- 5.42 Consider the decomposition of liquid benzene, $\text{C}_6\text{H}_6(l)$, to gaseous acetylene, $\text{C}_2\text{H}_2(g)$:
 $\frac{1}{3}\text{C}_6\text{H}_6(l) \longrightarrow \text{C}_2\text{H}_2(g) \quad \Delta H = +210 \text{ kJ}$
 (a) What is the enthalpy change for the reverse reaction? (b) What is ΔH for the decomposition of 1 mol of benzene to acetylene? (c) Which is more likely to be thermodynamically favored, the forward reaction or the reverse reaction? (d) If $\text{C}_6\text{H}_6(g)$ were consumed instead of $\text{C}_6\text{H}_6(l)$, would you expect the magnitude of ΔH to increase, decrease, or stay the same? Explain.

Calorimetry

- 5.43 (a) What are the units of heat capacity? (b) What are the units of specific heat?
- 5.44 Two solid objects, A and B, are placed in boiling water and allowed to come to temperature there. Each is then lifted out and placed in separate beakers containing 1000 g water at 10.0°C . Object A increases the water temperature by 3.50°C ; B increases the water temperature by 2.60°C . (a) Which object has the larger heat capacity? (b) What can you say about the specific heats of A and B?
- 5.45 (a) What is the specific heat of liquid water? (b) What is the heat capacity of 185 g of liquid water? (c) How many kJ of heat are needed to raise the temperature of 10.00 kg of liquid water from 24.6°C to 46.2°C ?
- 5.46 (a) What is the molar heat capacity of liquid water? (b) What is the heat capacity of 8.42 mol of liquid water? (c) How many kJ of heat are needed to raise the temperature of 2.56 kg of water from 44.8°C to 92.0°C ?
- 5.47 The specific heat of copper metal is $0.385 \text{ J/g}\cdot\text{K}$. How many J of heat are necessary to raise the temperature of a 1.42-kg block of copper from 25.0°C to 88.5°C ?
- 5.48 The specific heat of toluene (C_7H_8), is $1.13 \text{ J/g}\cdot\text{K}$. How many J of heat are needed to raise the temperature of 62.0 g of toluene from 16.3°C to 38.8°C ?
- 5.49 When a 9.55-g sample of solid sodium hydroxide dissolves in 100.0 g of water in a coffee-cup calorimeter (Figure 5.18), the temperature rises from 23.6°C to 47.4°C . Calculate ΔH (in kJ/mol NaOH) for the solution process
 $\text{NaOH}(s) \longrightarrow \text{Na}^+(aq) + \text{OH}^-(aq)$
 Assume that the specific heat of the solution is the same as that of pure water.
- 5.50 When a 3.88-g sample of solid ammonium nitrate dissolves in 60.0 g of water in a coffee-cup calorimeter (Figure 5.18), the temperature drops from 23.0°C to 18.4°C . Calculate ΔH (in kJ/mol NH_4NO_3) for the solution process
 $\text{NH}_4\text{NO}_3(s) \longrightarrow \text{NH}_4^+(aq) + \text{NO}_3^-(aq)$
 Assume that the specific heat of the solution is the same as that of pure water.
- 5.51 A 2.200-g sample of quinone ($\text{C}_6\text{H}_4\text{O}_2$) is burned in a bomb calorimeter whose total heat capacity is $7.854 \text{ kJ}/^\circ\text{C}$. The temperature of the calorimeter increases from 23.44°C to 30.57°C . What is the heat of combustion per gram of quinone? Per mole of quinone?
- 5.52 A 1.800-g sample of phenol ($\text{C}_6\text{H}_5\text{OH}$) was burned in a bomb calorimeter whose total heat capacity is $11.66 \text{ kJ}/^\circ\text{C}$. The temperature of the calorimeter plus contents increased from 21.36°C to 26.37°C . (a) Write a balanced chemical equation for the bomb calorimeter reaction. (b) What is the heat of combustion per gram of phenol? Per mole of phenol?

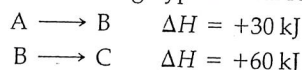
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- 5.53 Under constant-volume conditions the heat of combustion of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) is 15.57 kJ/g. A 2.500-g sample of glucose is burned in a bomb calorimeter. The temperature of the calorimeter increased from 20.55°C to 23.25°C. (a) What is the total heat capacity of the calorimeter? (b) If the size of the glucose sample had been exactly twice as large, what would the temperature change of the calorimeter have been?
- 5.54 Under constant-volume conditions the heat of combustion of benzoic acid ($\text{HC}_7\text{H}_5\text{O}_2$) is 26.38 kJ/g. A 1.640-g sample of benzoic acid is burned in a bomb calorimeter.

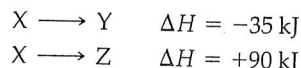
The temperature of the calorimeter increases from 22.25°C to 27.20°C. (a) What is the total heat capacity of the calorimeter? (b) A 1.320-g sample of a new organic substance is combusted in the same calorimeter. The temperature of the calorimeter increases from 22.14°C to 26.82°C. What is the heat of combustion per gram of the new substance? (c) Suppose that in changing samples, a portion of the water in the calorimeter were lost. In what way, if any, would this change the heat capacity of the calorimeter?

Hess's Law

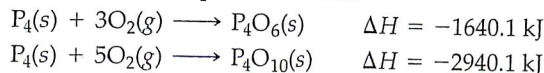
- 5.55 State Hess's law. Why is it important to thermochemistry?
- 5.56 What is the connection between Hess's law and the fact that H is a state function?
- 5.57 Consider the following hypothetical reactions:



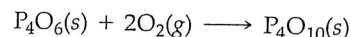
- (a) Use Hess's law to calculate the enthalpy change for the reaction $\text{A} \longrightarrow \text{C}$. (b) Construct an enthalpy diagram for substances A, B, and C, and show how Hess's law applies.
- 5.58 Suppose you are given the following hypothetical reactions:



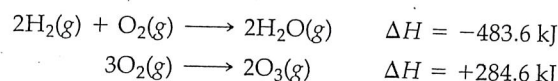
- (a) Use Hess's law to calculate the enthalpy change for the reaction $\text{Y} \longrightarrow \text{Z}$. (b) Construct an enthalpy diagram for substances X, Y, and Z. (c) Would it be valid to do what we have asked in part (a) if the first reaction had been carried out at 25°C and the second at 240°C? Explain.
- 5.59 Given the enthalpies of reaction



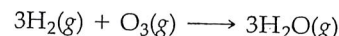
calculate the enthalpy change for the reaction



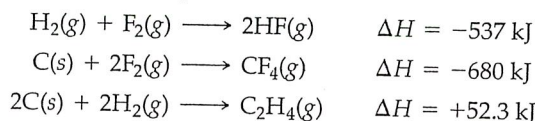
- 5.60 From the heats of reaction



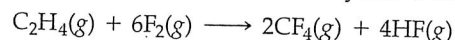
calculate the heat of the reaction



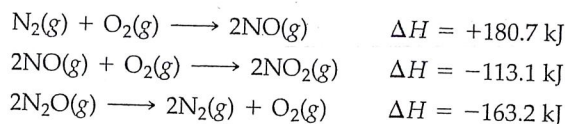
- 5.61 From the enthalpies of reaction



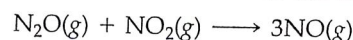
calculate ΔH for the reaction of ethylene with F_2 :



- 5.62 Given the data



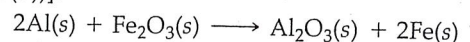
use Hess's law to calculate ΔH for the reaction



Enthalpies of Formation

- 5.63 (a) What is meant by the term *standard conditions*, with reference to enthalpy changes? (b) What is meant by the term *enthalpy of formation*? (c) What is meant by the term *standard enthalpy of formation*?
- 5.64 (a) Why are tables of standard enthalpies of formation so useful? (b) What is the value of the standard enthalpy of formation of an element in its most stable form?
- 5.65 Suppose it were decided that the standard enthalpies of formation of all elements in their most stable forms should be 100 kJ/mol. Would it still be possible to have standard enthalpies of formation of compounds, as in Table 5.3? If so, would any of the values in Table 5.3 be the same? Explain.
- 5.66 Using Table 5.3, determine whether the reaction of solid sucrose with liquid water to form solid glucose is an endothermic or exothermic process.

- 5.67 For each of the following compounds, write a balanced thermochemical equation depicting the formation of 1 mol of the compound from its elements in their standard states and use Appendix C to obtain the value of ΔH_f° : (a) $\text{NH}_3(\text{g})$; (b) $\text{SO}_2(\text{g})$; (c) $\text{RbClO}_3(\text{s})$; (d) $\text{NH}_4\text{NO}_3(\text{s})$.
- 5.68 Write balanced equations that describe the formation of the following compounds from their elements in their standard states, and use Appendix C to obtain the values of their standard enthalpies of formation: (a) $\text{HBr}(\text{g})$; (b) $\text{AgNO}_3(\text{s})$; (c) $\text{Hg}_2\text{Cl}_2(\text{s})$; (d) $\text{C}_2\text{H}_5\text{OH}(\text{l})$.
- 5.69 The following is known as the thermite reaction [(Figure 5.8(b))]:



This highly exothermic reaction is used for welding massive units, such as propellers for large ships. Using enthalpies of formation in Appendix C, calculate ΔH° for this reaction.

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- 5.70 Many cigarette lighters contain liquid butane, $C_4H_{10}(l)$. Using enthalpies of formation, calculate the quantity of heat produced when 1.0 g of butane is completely combusted in air.
- 5.71 Using values from Appendix C, calculate the standard enthalpy change for each of the following reactions:
- $2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$
 - $Mg(OH)_2(s) \longrightarrow MgO(s) + H_2O(l)$
 - $4FeO(s) + O_2(g) \longrightarrow 2Fe_2O_3(s)$
 - $SiCl_4(l) + 2H_2O(l) \longrightarrow SiO_2(s) + 4HCl(g)$
- 5.72 Using values from Appendix C, calculate the value of ΔH° for each of the following reactions:
- $N_2O_4(g) + 4H_2(g) \longrightarrow N_2(g) + 4H_2O(g)$
 - $2KOH(s) + CO_2(g) \longrightarrow K_2CO_3(s) + H_2O(g)$
 - $SO_2(g) + 2H_2S(g) \longrightarrow (\frac{3}{8})S_8(s) + 2H_2O(g)$
 - $Fe_2O_3(s) + 6HCl(g) \longrightarrow 2FeCl_3(s) + 3H_2O(g)$
- 5.73 Complete combustion of 1 mol of acetone (C_3H_6O) liberates 1790 kJ:
- $$C_3H_6O(l) + 4O_2(g) \longrightarrow 3CO_2(g) + 3H_2O(l) \quad \Delta H^\circ = -1790 \text{ kJ}$$
- Using this information together with data from Appendix C, calculate the enthalpy of formation of acetone.
- 5.74 Calcium carbide (CaC_2) reacts with water to form acetylene (C_2H_2) and $Ca(OH)_2$. From the following enthalpy of reaction data and data in Appendix C, calculate ΔH_f° for $CaC_2(s)$:
- $$CaC_2(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(s) + C_2H_2(g) \quad \Delta H^\circ = -127.2 \text{ kJ}$$
- 5.75 Calculate the standard enthalpy of formation of solid $Mg(OH)_2$, given the following data:
- $$2Mg(s) + O_2(g) \longrightarrow 2MgO(s) \quad \Delta H^\circ = -1203.6 \text{ kJ}$$
- $$Mg(OH)_2(s) \longrightarrow MgO(s) + H_2O(l) \quad \Delta H^\circ = +37.1 \text{ kJ}$$
- $$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l) \quad \Delta H^\circ = -571.7 \text{ kJ}$$
- 5.76 (a) Calculate the standard enthalpy of formation of gaseous diborane (B_2H_6) using the following thermodynamic information:
- $$4B(s) + 3O_2(g) \longrightarrow 2B_2O_3(s) \quad \Delta H^\circ = -2509.1 \text{ kJ}$$
- $$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l) \quad \Delta H^\circ = -571.7 \text{ kJ}$$
- $$B_2H_6(g) + 3O_2(g) \longrightarrow B_2O_3(s) + 3H_2O(l) \quad \Delta H^\circ = -2147.5 \text{ kJ}$$
- (b) Pentaborane (B_5H_9) is another in a series of boron hydrides. What experiment or experiments would you need to perform to yield the data necessary to calculate the heat of formation of $B_5H_9(l)$? Explain by writing out and summing any applicable chemical reactions.
- 5.77 Gasoline is composed primarily of hydrocarbons, including many with eight carbon atoms, called *octanes*. One of the cleanest burning octanes is a compound called 2,3,4-trimethylpentane, which has the following structural formula:
- $$\begin{array}{ccccccc} & CH_3 & & CH_3 & & CH_3 & \\ & | & & | & & | & \\ H_3C & -CH & - & CH & - & CH & -CH_3 \end{array}$$
- The complete combustion of 1 mol of this compound to $CO_2(g)$ and $H_2O(g)$ leads to $\Delta H^\circ = -5069 \text{ kJ}$. (a) Write a balanced equation for the combustion of 1 mol of $C_8H_{18}(l)$. (b) Write a balanced equation for the formation of $C_8H_{18}(l)$ from its elements. (c) By using the information in this problem and data in Table 5.3, calculate ΔH_f° for 2,3,4-trimethylpentane.
- 5.78 Naphthalene ($C_{10}H_8$) is a solid aromatic compound often sold as mothballs. The complete combustion of this substance to yield $CO_2(g)$ and $H_2O(l)$ at $25^\circ C$ yields 5154 kJ/mol. (a) Write balanced equations for the formation of naphthalene from the elements and for its combustion. (b) Calculate the standard enthalpy of formation of naphthalene.

Foods and Fuels

- 5.79 (a) What is meant by the term *fuel value*? (b) What substance is often referred to as *blood sugar*? Why is it significant in the discussion of human foods? (c) Which is a greater source of energy as food, 5 g of fat or 9 g of carbohydrate?
- 5.80 (a) Why are fats well suited for energy storage in the human body? (b) A particular chip snack food is composed of 12% protein, 14% fat, and the rest carbohydrate. What percentage of the calorie content of this food is fat? (c) How many grams of protein provide the same fuel value as 25 g of fat?
- 5.81 A serving of Campbell's® condensed cream of mushroom soup contains 7 g fat, 9 g carbohydrate, and 1 g protein. Estimate the number of Calories in a serving.
- 5.82 A pound of plain M&M® candies contains 96 g fat, 320 g carbohydrate, and 21 g protein. What is the fuel value in kJ in a 42-g (about 1.5 oz) serving? How many Calories does it provide?
- 5.83 The heat of combustion of fructose, $C_6H_{12}O_6$, is -2812 kJ/mol . If a fresh golden delicious apple weighing 4.23 oz (120 g) contains 16.0 g of fructose, what caloric content does the fructose contribute to the apple?
- 5.84 The heat of combustion of ethanol, $C_2H_5OH(l)$, is -1367 kJ/mol . A batch of sauvignon blanc wine contains 10.6% ethanol by mass. Assuming the density of the wine to be 1.0 g/mL, what caloric content does the alcohol (ethanol) in a 6-oz glass of wine (177 mL) have?
- 5.85 The standard enthalpies of formation of gaseous propyne (C_3H_4), propylene (C_3H_6), and propane (C_3H_8) are $+185.4$, $+20.4$, and -103.8 kJ/mol , respectively. (a) Calculate the heat evolved per mole on combustion of each substance to yield $CO_2(g)$ and $H_2O(g)$. (b) Calculate the heat evolved on combustion of 1 kg of

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Name: _____ Class: _____

General Chapter 5 Review

Thermo

This is a general review for the material covered in this chapter. It is intended to be a supplementary study tool that should be used in conjunction with all notes, worksheets, and your text. Please show all work and circle your final answer for full credit. (88 PTS.)

1. Define the following:

a. Kinetic energy:

b. Formula for kinetic energy (identify the variables and their units):

c. 1 calorie = _____ Joule

d. Potential energy

e. State Function

f. Specific heat

g. Heat capacity

h. Calorimeter

2. Draw a picture of an exothermic system having work done on it by the surroundings. Label system, surroundings, heat, and work.

3. What is the first law of thermodynamics?

4. A system releases 437.00 kJ of heat and 467.00 J of work are done on the system. What is the internal energy of the process? Is this process exothermic or endothermic?

5. Circle the following that **ARE** state functions

Heat

Work

Internal energy

Potential energy

Kinetic energy

Enthalpy

Time

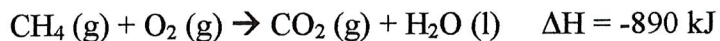
Temperature

6. What form of work is present in the thermodynamic function of enthalpy?

7. Draw an enthalpy diagram for an exothermic reaction, include the ΔH value.

8. Draw an enthalpy diagram for an endothermic reaction, include the ΔH value.

9. How much heat is released when 6.50 g of oxygen gas is burned in a constant pressure system according to the equation below? BECAREFUL!



10. How much heat is needed to warm 250.0 g of water from 22.0 °C to 98.0 °C.

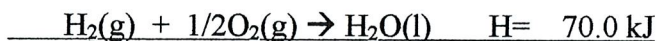
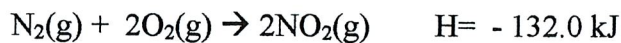
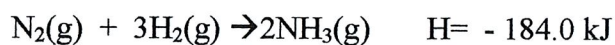
11. The specific heat of calcium is 0.179 J/g-K. How many J of heat are necessary to raise the temperature of 1.42 kg block of calcium from 35.0 °C to 98.0 °C?

12. The specific heat of benzene is 2.34 J/g-K. How many Joules of heat are needed to raise the temperature of 63.0 g of benzene from 16.0 °C to 38.0 °C?

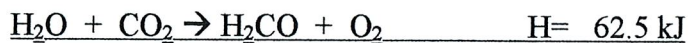
13. When 4.04 g of solid ammonium nitrate dissolve in 60.0g of water in a coffee cup calorimeter. The temperature of the water drops from 23.0°C to 16.0 °C. Calculate change in enthalpy for the solution in kJ/mol of NH_4NO_3 .



15. Use the thermochemical equations shown below to determine the enthalpy for the reaction:



16. Use the thermochemical equations shown below to determine the enthalpy for the reaction:



Use table appendix 3 to answer the following questions.

17. Identify the standard state (solid, liquid or gas) for the following elements:

Place an X in the box if a value is not given.

UNITS _____

	<u>***Solid***</u>	<u>Liquid- aq</u>	<u>Gas</u>
Carbon dioxide			
sodium			
water			
KOH			
Sulfuric Acid			

18. Calculate the ΔH° for the reaction:

