THE FLOW OF ENERGY—
HEAT AND WORK

Section Review

Objectives
- Explain the relationship between energy, heat, and work
- Distinguish between exothermic and endothermic processes
- Distinguish between heat capacity and specific heat

Vocabulary
- thermochemistry
- chemical potential energy
- heat
- system
- surroundings
- law of conservation of energy
- endothermic process
- exothermic process
- heat capacity
- specific heat

Key Equations and Relationships
- 1 Calorie = 1 kilocalorie = 1000 calories
- 1 J = 0.2390 cal and 4.184 J = 1 cal
- \[ C = \frac{q}{m \times \Delta T} = \frac{\text{heat (joules or calories)}}{\text{mass (g)} \times \text{change in temperature (°C)}} \]

Part A Completion

Use this completion exercise to check your understanding of the concepts and terms that are introduced in this section. Each blank can be completed with a term, short phrase, or number.

The energy that flows from a warm object to a cool object is called __1___. The energy stored within the structural units of chemical substances is called chemical __2___. The study of heat transfer during chemical reactions and changes of state is called __3___. One of the units used to measure heat flow is the __4____, defined as the amount of heat needed to raise 1 g of water 1°C. The SI unit of heat and energy is the __5____, which is equal to 0.2390 cal. The __6____ of a substance is the amount of heat it takes to change the temperature of 1 g of the substance 1°C. Substances like __7____, with low heat capacities, take a shorter time to heat up than substances with high heat capacities, such as __8___.

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Part B True-False
Classify each of these statements as always true, AT; sometimes true, ST; or never true, NT.

9. The joule is the SI unit of force.  AT
10. Endothermic processes absorb heat from the surroundings.  NT
11. The law of conservation of energy states that in a chemical process, energy is sometimes created and sometimes destroyed.  AT
12. A system that loses heat to its surroundings is said to be exothermic, and the value of \( q \) is negative.  AT
13. A calorie is defined as the quantity of heat needed to raise the temperature of 1 gram of pure water 1°C.  AT

Part C Matching
Match each description in Column B to the correct term in Column A.

<table>
<thead>
<tr>
<th>Column A</th>
<th>Column B</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>a. a process that absorbs heat from the surroundings</td>
</tr>
<tr>
<td>E</td>
<td>b. the amount of heat required to change the temperature of an object by exactly 1°C</td>
</tr>
<tr>
<td>B</td>
<td>c. energy that transfers from one object to another because of a temperature difference between them</td>
</tr>
<tr>
<td>D</td>
<td>d. the part of the universe being studied</td>
</tr>
<tr>
<td>A</td>
<td>e. a process that loses heat to the surroundings</td>
</tr>
</tbody>
</table>

Part D Questions and Problems
Answer the following in the space provided.

19. Distinguish among the various forms of energy: chemical potential energy, work, and heat.

\[
\text{Energy stored in chemical bonds} \\
\text{Force through a distance} \\
\text{Energy transfer between objects at different temperatures}
\]

20. The temperature of a piece of unknown metal with a mass of 18.0 g increases from 25.0°C to 40°C when the metal absorbs 124.2 J of heat. What is the specific heat of the unknown metal? Compare your answer to the values listed in Table 17.2 of your textbook. What is the identity of the unknown metal?

\[
q = \Delta \text{H} = \text{m} \cdot c \cdot \Delta T \\
124.2 \text{ J} = 18.0 \text{ g} \cdot c \cdot (40 - 25) \text{ } \text{C} \\
c = \frac{124.2 \text{ J}}{18.0 \text{ g} \cdot 15 \text{ C}} \approx 0.465 \text{ J/g°C}
\]
MEASURING AND EXPRESSING
ENTHALPY CHANGES

Section Review

Objectives
• Construct equations that show the enthalpy changes for chemical and physical processes
• Calculate enthalpy changes in chemical and physical processes

Vocabulary
• calorimetry
• calorimeter
• enthalpy, \( (H) \)
• thermochemical equation
• heat of reaction
• heat of combustion

Key Equation
\[ q_{\text{sys}} = \Delta H = -q_{\text{arr}} = -m \times c \times \Delta T, \text{where } \Delta T = T_f - T_i \]

Part A Completion
Use this completion exercise to check your understanding of the concepts and terms that are introduced in this section. Each blank can be completed with a term, short phrase, or number.

A ___1___ is a device used to measure the absorption or release of heat in chemical and physical processes. For systems at constant pressure, the heat changes that occur are the same as changes in ___2___, symbolized as ___3___. To measure the enthalpy change for a reaction in aqueous solution, it is necessary to measure the ___4___ and ___5___ temperatures of the system and the ___6___ of the water in the system.

1. Calorimeter
2. Enthalpy
3. \( \Delta H \)
4. Initial
5. Final
6. Mass

Part B True-False
Classify each of these statements as always true, AT; sometimes true, ST; or never true, NT.

7. When a substance dissolves in water, heat is released.  ST
8. The sign of \( \Delta H \) is negative for an exothermic reaction.  AT
9. If 129 kJ of heat is required to decompose 2 moles of NaHCO₃, then 258 kJ is required to decompose 4 moles of NaHCO₃.  AT
10. The physical state of the reactants and products in a thermochemical reaction are not important when calculating ΔH of the reaction.

11. In endothermic reactions, the potential energy of the product(s) is higher than the potential energy of the reactants.

12. The equation CaO(s) + H₂O(l) → Ca(OH)₂(s)  ΔH = 65.2 kJ is an example of a thermochemical equation.

**Part C Matching**

Match each description in Column B to the correct term in Column A.

**Column A**
- E
- A
- B
- C
- D

**Column B**
- 13. enthalpy (H)
- 14. heat of combustion
- 15. thermochemical equation
- 16. calorimetry
- 17. bomb calorimeter

a. the heat of reaction for the complete burning of 1 mole of a substance

b. a chemical equation that includes the enthalpy change (ΔH)

c. the accurate and precise measurement of heat changes for chemical and physical processes

d. an insulated device containing a sealed vessel that is used to measure the heat released during a combustion reaction

e. the amount of heat that a system has at a constant pressure

**Part D Questions and Problems**

Answer the following in the space provided.

18. When 2 moles of nitric oxide, NO, burn in air to produce 2 moles of nitrogen dioxide, 113.04 kJ of heat is produced. Write a balanced thermochemical equation for this reaction.

\[ 2 \text{NO} + \text{O}_2 \rightarrow 2 \text{NO}_2 \quad \Delta H = 113.04 \text{ kJ} \]

19. Calculate the amount of heat produced when 34.8 g of methane, CH₄, burns in an excess of oxygen, according to the following equation.

\[ \text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(l) \quad \Delta H = -890.2 \text{ kJ} \]

\[ 34.8 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16 \text{ g}} \times \frac{-890.2 \text{ kJ}}{1 \text{ mol CH}_4} = 1.94 \times 10^5 \text{ kJ} \]
Table 17.4
Standard Heats of Formation ($\Delta H^0$) at 25°C and 101.3 kPa

<table>
<thead>
<tr>
<th>Substance</th>
<th>$\Delta H^0$ (kJ/mol)</th>
<th>Substance</th>
<th>$\Delta H^0$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{Al}_2\text{O}_3(s)$</td>
<td>$-1676.0$</td>
<td>$\text{H}_2\text{O}_2(l)$</td>
<td>$-187.8$</td>
</tr>
<tr>
<td>$\text{Br}_2(g)$</td>
<td>$30.91$</td>
<td>$\text{I}_2(g)$</td>
<td>$62.4$</td>
</tr>
<tr>
<td>$\text{Br}_2(l)$</td>
<td>$0.0$</td>
<td>$\text{I}_2(s)$</td>
<td>$0.0$</td>
</tr>
<tr>
<td>$\text{C}(s, \text{diamond})$</td>
<td>$1.9$</td>
<td>$\text{N}_2(g)$</td>
<td>$0.0$</td>
</tr>
<tr>
<td>$\text{C}(s, \text{graphite})$</td>
<td>$0.0$</td>
<td>$\text{NH}_3(g)$</td>
<td>$-46.19$</td>
</tr>
<tr>
<td>$\text{CH}_4(g)$</td>
<td>$-74.86$</td>
<td>$\text{NO}(g)$</td>
<td>$90.37$</td>
</tr>
<tr>
<td>$\text{CO}(g)$</td>
<td>$-110.5$</td>
<td>$\text{NO}_2(g)$</td>
<td>$33.85$</td>
</tr>
<tr>
<td>$\text{CO}_2(g)$</td>
<td>$-393.5$</td>
<td>$\text{NaCl}(s)$</td>
<td>$-411.2$</td>
</tr>
<tr>
<td>$\text{CaCO}_3(s)$</td>
<td>$-1207.0$</td>
<td>$\text{O}_2(g)$</td>
<td>$0.0$</td>
</tr>
<tr>
<td>$\text{CaO}(s)$</td>
<td>$-635.1$</td>
<td>$\text{O}_3(g)$</td>
<td>$142.0$</td>
</tr>
<tr>
<td>$\text{Cl}_2(g)$</td>
<td>$0.0$</td>
<td>$\text{P}(s, \text{white})$</td>
<td>$0.0$</td>
</tr>
<tr>
<td>$\text{Fe}(s)$</td>
<td>$0.0$</td>
<td>$\text{P}(s, \text{red})$</td>
<td>$-18.4$</td>
</tr>
<tr>
<td>$\text{Fe}_2\text{O}_3(s)$</td>
<td>$-822.1$</td>
<td>$\text{S}(s, \text{rhombic})$</td>
<td>$0.0$</td>
</tr>
<tr>
<td>$\text{H}_2(g)$</td>
<td>$0.0$</td>
<td>$\text{S}(s, \text{monoclinic})$</td>
<td>$0.30$</td>
</tr>
<tr>
<td>$\text{H}_2\text{O}(g)$</td>
<td>$-241.8$</td>
<td>$\text{SO}_2(g)$</td>
<td>$-296.8$</td>
</tr>
<tr>
<td>$\text{H}_2\text{O}(l)$</td>
<td>$-285.8$</td>
<td>$\text{SO}_3(g)$</td>
<td>$-395.7$</td>
</tr>
</tbody>
</table>

**Standard Heats of Formation**

Enthalpy changes generally depend on conditions of the process. In order to compare enthalpy changes, scientists specify a common set of conditions as a reference point. These conditions, called the standard state, refer to the stable form of a substance at 25°C and 101.3 kPa. The **standard heat of formation** ($\Delta H^0_f$) of a compound is the change in enthalpy that accompanies the formation of one mole of a compound from its elements with all substances in their standard state at 25°C. The $\Delta H^0_f$ of a free element in its standard state is arbitrarily set at zero. Thus, $\Delta H^0_f = 0$ for the diatomic molecules $\text{H}_2(g)$, $\text{N}_2(g)$, $\text{O}_2(g)$, $\text{F}_2(g)$, $\text{Cl}_2(g)$, $\text{Br}_2(l)$, and $\text{I}_2(s)$. Similarly, $\Delta H^0_f = 0$ for the graphite form of carbon, $\text{C}(s, \text{graphite})$. Table 17.4 lists $\Delta H^0_f$ values for some common substances.

Standard heats of formation provide an alternative to Hess's law in determining heats of reaction indirectly. **For a reaction that occurs at standard conditions, you can calculate the heat of reaction by using standard heats of formation.** Such an enthalpy change is called the standard heat of reaction ($\Delta H^0_r$). The standard heat of reaction is the difference between the standard heats of formation of all the reactants and products.

$$\Delta H^0_r = \Delta H^0_f(\text{products}) - \Delta H^0_f(\text{reactants})$$

Figure 17.15 is an enthalpy diagram for the formation of water from its elements at standard conditions. The enthalpy difference between the reactants and products, $-285.8 \text{ kJ/mol}$, is the standard heat of formation of liquid water from the gases hydrogen and oxygen. Notice that water has a lower enthalpy than the elements from which it is formed.
CALCULATING HEATS OF REACTION

Section Review

Objectives

- Apply Hess’s law of heat summation to find enthalpy changes for chemical and physical processes
- Calculate enthalpy changes using standard heats of formation

Vocabulary

- Hess’s law of heat summation
- standard heat of formation

Key Equation

- $\Delta H^\circ = \Delta H^\circ_{\text{products}} - \Delta H^\circ_{\text{reactants}}$

Part A Completion

Use this completion exercise to check your understanding of the concepts and terms that are introduced in this section. Each blank can be completed with a term, short phrase, or number.

Hess’s law of heat summation states that for a chemical equation that can be written as the ___1___ of two or more steps, the ___2___ change for the final equation equals the sum of the enthalpy changes for the individual steps. Hess’s law makes it possible to measure the heat of a reaction ___3___. When a reaction is reversed, the sign of $\Delta H$ must be ___4___.

Sometimes it is hard to measure the heat for a reaction. In such cases, the ___5___ is used to calculate heats of reaction at standard conditions. The standard heat of formation of a compound is the ___6___ in enthalpy that accompanies the formation of ___7___ mole of a compound from its elements.

The symbol used for standard heat of formation is ___8___. The standard heat of formation of a free element in its standard state is ___9___. The standard heat of reaction is determined by ___10___ the $\Delta H^\circ$ of all the reactants from the $\Delta H^\circ$ of all the products.

1. sum
2. Heat
3. $\Delta H^\circ_{\text{rxn}}$
4. reversed
5. equation
6. change
7. one
8. $\Delta H^\circ$
9. zero
10. difference between
Part B True-False

Classify each of these statements as always true, AT; sometimes true, ST; or never true, NT.

11. The standard heat of formation for a substance is determined at 100°C. **NT**
12. Hess's law of heat summation is not related to the law of conservation of energy. **AT**
13. When using Hess's law of heat summation, intermediate reactions are summed and terms are canceled, as in algebra, to arrive at a final equation. **UT**
14. The $\Delta H_f$ for $I_2(g)$ is zero. **UT**
15. The $\Delta H_f$ for $H_2O(l)$ and $H_2O(s)$ are the same. **UT**

Part C Matching

Match each description in Column B to the correct term in Column A.

<table>
<thead>
<tr>
<th>Column A</th>
<th>Column B</th>
</tr>
</thead>
<tbody>
<tr>
<td>B 16. standard heat of formation</td>
<td>a. symbol for the standard heat of formation</td>
</tr>
<tr>
<td>C 17. Hess's law of heat summation</td>
<td>b. the change in enthalpy that accompanies the formation of 1 mole of a compound from its elements</td>
</tr>
<tr>
<td>A 18. $\Delta H_f$</td>
<td>c. in going from a particular set of reactants to a particular set of products, the enthalpy change is the same whether the reaction takes place in one step or in a series of steps</td>
</tr>
<tr>
<td>D 19. $\Delta H_f$ for $Br_2(g) \rightarrow Br_2(l)$</td>
<td>d. $-30.91$ kJ</td>
</tr>
<tr>
<td>E 20. zero</td>
<td>e. $\Delta H_f$ of $Cl_2(g)$</td>
</tr>
</tbody>
</table>

Part D Questions and Problems

Answer the following in the space provided.

21. Determine the heat of reaction for the following reaction.

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$$

Use the following thermochemical equations.

1) $CuO(s) \rightarrow Cu(s) + \frac{1}{2}O_2(g)$ \hspace{1cm} $\Delta H = 155$ kJ
2) $H_2O(g) \rightarrow H_2(g) + \frac{1}{2}O_2(g)$ \hspace{1cm} $\Delta H = -242$ kJ

$$\Delta H_{rxn} = -87.5$$

22. Calculate the change in enthalpy for the following reaction using standard heats of formation. (Refer to Table 17.4 in your textbook.)

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

$$\Delta H_{rxn} =$$
2) \[ \text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2 \]

\[ -822.2 \text{kJ} - 110.5 \text{kJ} + 0 - 395.5 \text{kJ} \]

\[ \Delta H^\circ_{\text{rxn}} = [3(-395.5 \text{kJ})] - \]

\[ -822.2 \text{kJ} + 3(-110.5 \text{kJ}) \]

\[ = -27 \text{kJ} \]
Practice Problems
In your notebook, solve the following problems.

SECTION 17.1 THE FLOW OF ENERGY—HEAT AND WORK

Use the three-step problem-solving approach you learned in Chapter 1.

1. How many kilojoules of energy are in a donut that contains 200.0 Calories?

2. What is the specific heat of a substance that has a mass of 25.0 g and requires 525.0 calories to raise its temperature by 15.0°C?

3. Suppose 100.0 g of H₂O(s) absorbs 1255.0 J of heat. What is the corresponding temperature change? The specific heat capacity of H₂O(s) is 2.1 J/g·°C.

4. How many joules of heat energy are required to raise the temperature of 100.0 g of aluminum by 120.0°C? The specific heat capacity of aluminum is 0.90 J/g·°C.

SECTION 17.2 MEASURING AND EXPRESSING ENTHALPY CHANGES

1. A student mixed 75.0 mL of water containing 0.75 mol HCl at 25°C with 75.0 mL of water containing 0.75 mol of NaOH at 25°C in a foam cup calorimeter. The temperature of the resulting solution increased to 35°C. How much heat in kilojoules was released by this reaction?

   \[ C_{\text{water}} = 4.18 \text{ J/g·°C} \]

2. Calculate the amount of heat evolved when 15.0 g of Ca(OH)₂ forms from the reaction of CaO(s) + H₂O(l).

   \[ \text{CaO(s) + H}_2\text{O(l) }\rightarrow \text{Ca(OH)}_2(s) \quad \Delta H = -65.2 \text{ kJ} \]

3. Calculate the amount of heat produced when 52.4 g of methane, CH₄, burns in an excess of air, according to the following equation.

   \[ \text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O(l)} \quad \Delta H = -890.2 \text{ kJ} \]

4. Balance the following equation, then calculate the enthalpy change for the reaction given that the standard heat of combustion of NH₃(g) is -226 kJ/mol.

   \[ \text{NH}_3(g) + \text{O}_2(g) \rightarrow \text{NO(g) + H}_2\text{O(g)} \]

SECTION 17.3 HEAT IN CHANGES OF STATE

1. Calculate the amount of heat needed to melt 35.0 g of ice at 0°C. Express your answer in kilojoules.

2. Calculate the amount of heat needed to convert 190.0 g of liquid water at 18°C to steam at 100.0°C.

3. How much heat (kJ) is released when 2.543 mol NaOH(s) is dissolved in water?

   \[ \text{NaOH(s)} \overset{\text{H}_2\text{O(l)}}{\rightarrow} \text{Na}^+(aq) + \text{OH}^-(aq) \quad \Delta H_{\text{soln}} = -445.1 \text{ kJ/mol} \]
4. Calculate the amount of heat needed to convert 96 g of ice at -24°C to water at 28°C. The specific heat capacity of H₂O(s) is 2.1 J/g°C.

**SECTION 17.4 CALCULATING HEATS OF REACTION**

1. What is the standard heat of reaction for the combustion of hydrogen sulfide? Refer to Table 17.4 in your textbook.
   \[ 2\text{H}_2\text{S}(g) + 3\text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g) + 2\text{SO}_2(g) \]
   \[ \Delta H = -296.1 \text{ kJ} \]

2. Calculate the enthalpy change (in kJ) for the following reaction. State whether the reaction is exothermic or endothermic. Refer to Table 17.4 in your textbook.
   \[ \text{CaO}(s) + \text{CO}_2(g) \rightarrow \text{CaCO}_3(s) \]
   \[ -636.6 \text{ kJ} \]

3. What is the enthalpy change for the formation of hydrazine, N₂H₄(l), from its elements?
   \[ \text{N}_2(g) + 2\text{H}_2(g) \rightarrow \text{N}_2\text{H}_4(l) \]
   \[ \Delta H = -393.5 \text{ kJ} \]

   Use the following reactions and enthalpy changes:
   \[ \text{N}_2\text{H}_4(l) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{N}_2(g) + 2\text{H}_2\text{O}(l) \]
   \[ \Delta H = -622.2 \text{ kJ} \]
   \[ \text{H}_2(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{H}_2\text{O}(l) \]
   \[ \Delta H = -285.8 \text{ kJ} \]

   \[ \text{N}_2 + 2\text{H}_2\text{O} \rightarrow \text{N}_2\text{H}_4 + \frac{3}{2}\text{O}_2 \]
   \[ +622.2 \text{ kJ} \]

   \[ 2\text{H}_2 + \frac{3}{2}\text{O}_2 \rightarrow 2\text{H}_2\text{O} \]
   \[ 2(-285.8) \]
   \[ +506 \text{ kJ} \]

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438 Core Teaching Resources
17.1

#1) \[ 200.0 \text{cal} \times \frac{4.184 \text{J}}{1 \text{cal}} = 836.8 \text{J} \]

#4) \[ \ell = (1000 \text{g})(0.86 \frac{\text{J}}{\text{g} \cdot \text{°C}})(120 \text{°C}) \]

\[ = 10800 \text{J} \times \frac{1 \text{kJ}}{1000 \text{J}} = 10.8 \text{kJ} \]

17.2

#2) \[ \text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 \]
\[ \Delta H^\circ = -65.2 \text{kJ} \]

\[ 15 \text{g} \text{Ca(OH)}_2 \times \frac{1 \text{mol}}{74 \text{g}} \times \frac{-65.2 \text{kJ}}{1 \text{mol}} \]
\[ = -13.2 \text{kJ} \]

#3) \[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \]
\[ \Delta H^\circ = -890.2 \text{kJ} \]

\[ 54 \text{g} \text{CH}_4 \times \frac{1 \text{mol}}{16 \text{g}} \times \frac{-890.2 \text{kJ}}{1 \text{mol}} \]
\[ = -3004 \text{kJ} \]
\[
\begin{align*}
\frac{17.4}{\text{#1}} & \quad 2 \text{H}_2 \text{S}_8 \text{O}_3(\ell) + 3 \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(\ell) + 2 \text{SO}_2(g) \\
& \quad -20.15 \text{kJ/mol} \quad -296.1 \text{kJ/mol} \\
& \quad -241.8 \text{kJ/mol}
\end{align*}
\]

\[
\left[ 2(-241.8 \text{kJ}) + 2(-296.1 \text{kJ}) \right] - \left[ 2(-20.15 \text{kJ}) + 0 \right]
\]

\[
\Delta H_{f}^{\circ} = -1056 \text{kJ}
\]
17.4

\[ \text{CaO(s)} + \text{CO}_2(g) \rightarrow \text{CaCO}_3(s) \]

\[ \Delta H^{\circ} = [\text{-1206.9 KJ}] - \]

[\text{-635.6 KJ + 393.5 KJ}]

= -178 KJ

\#3

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

\[ \Delta H^{\circ} = +622.2 \text{ KJ} \]

\[ 2\text{H}_2 + \frac{3}{2}\text{O}_2 \rightarrow 2\text{H}_2\text{O} \]

\[ \Delta H^{\circ} = 2(-285.8 \text{ KJ}) \]

\[ \text{N}_2 + 2\text{H}_2 \rightarrow \text{N}_2\text{H}_4 \]

\[ \Delta H^{\circ} = +50.6 \text{ KJ} \]
Calorimetry Problems

$q_{\text{sur}} = m \times C \times \Delta T$
$q_{\text{run}} = -q_{\text{sur}}$

$q = \text{heat} \quad m = \text{mass} \quad \Delta T = T_f - T_i \quad C = \text{specific heat} \ (\text{for water} = 4.184 \ J/\text{g}^\circ \text{C})$

1. What is the specific heat of aluminum if the temperature of a 28.4 g sample of aluminum is increased by 8.1 °C when 207 J of heat is added?

\[
C = \frac{q}{m \Delta T} = \frac{207 \ J}{28.4 \ g \times 8.1 ^\circ C} = 0.907 \ \text{J/g}^\circ \text{C}
\]

2. What is the specific heat of silicon if the temperature of a 4.11 g sample of silicon is increased by 3.8 °C when 11.1 J of heat is added?

\[
C = \frac{q}{m \Delta T} = \frac{11.1 \ J}{4.11 \ g \times 3.8 ^\circ C} = 0.715 \ \text{J/g}^\circ \text{C}
\]

3. How much heat must be added to a 8.21 g sample of gold to increase its temperature by 6.2 °C? The specific heat of gold is 0.13 J/g°C.

\[
q = m \times C \times \Delta T = 8.21 \ g \times 0.13 \ \text{J/g}^\circ \text{C} \times 6.2 ^\circ C = 6.12 \ J
\]

4. If 40.5 J of heat is added to a 15.4 g sample of silver, how much will the temperature increase by? The specific heat of silver is 0.235 J/g°C.

\[
\Delta T = \frac{q}{m \times C} = \frac{40.5 \ J}{15.4 \ g \times 0.235 \ \text{J/g}^\circ \text{C}} = 9.97 \ ^\circ \text{C}
\]

5. When a 25.7 g sample of NaI dissolves in 80.0 g of water in a calorimeter, the temperature rises from 20.5 °C to 24.4 °C. Calculate ΔH for the process.

\[
\text{NaI(s)} \rightarrow \text{Na}^+(aq) + \text{I}^-(aq) \quad \Delta H = ?
\]

\[
\Delta T = \frac{q}{m \times C} = \frac{24.1 - 20.5}{80.0 \ g \times 4.184 \ \text{J/g}^\circ \text{C}} = 0.9 \ ^\circ \text{C}
\]

6. When a 16.9 g sample of NaOH dissolves in 70.0 g of water in a calorimeter, the temperature rises from 22.4 °C to 86.6 °C. Calculate ΔH for the process.

\[
\text{NaOH(s)} \rightarrow \text{Na}^+(aq) + \text{OH}^-\text{(aq)} \quad \Delta H = ?
\]

\[
\Delta T = \frac{q}{m \times C} = \frac{86.6 - 22.4}{70.0 \ g \times 4.184 \ \text{J/g}^\circ \text{C}} = 1.38 \ ^\circ \text{C}
\]
7. When a 12.8 g sample of KCl dissolves in 75.0 g of water in a calorimeter, the temperature drops from 31.0 °C to 21.6 °C. Calculate ΔH for the process.

\[
\text{KCl(s) } \rightarrow \text{ K}^+(aq) + \text{ Cl}^-(aq) \quad \Delta H = ?
\]

8. A 2.50 g sample of zinc is heated, then placed in a calorimeter containing 65.0 g of water. Temperature of water increases from 20.00 °C to 22.50 °C. The specific heat of zinc is 0.390 J/g°C. What was the initial temperature of the zinc metal sample? (final temperatures of zinc and water are the same)

9. A 13.5 g sample of gold is heated, then placed in a calorimeter containing 60.0 g of water. Temperature of water increases from 19.00 °C to 20.00 °C. The specific heat of gold is 0.130 J/g°C. What was the initial temperature of the gold metal sample?

\[
\left(13.50\right)\left(0.135\right)\left(20 - x\right) = \left(60.0\right)\left(4.184\right)\left(19 - 20\right)
\]

\[
1.755\left(x - 20\right) = -251
\]

\[
1.755x - 35.1 = -251
\]
1) \[ Z = \eta C \Delta T \]

\[ 2075 = (28.4 \, \text{kJ})(C)(8.1 \, ^\circ C) \]

\[ C = \frac{0.90 \, \text{J}}{\text{g} \, ^\circ C} \]

2) \[ 11.1 \, \text{J} = (4.11 \, \text{kJ})(C)(3.8 \, ^\circ C) \]

\[ C = \frac{0.71 \, \text{J}}{\text{g} \, ^\circ C} \]

3) \[ Z = (8.2 \, \text{kJ})(0.13 \, \text{J}) \left( \frac{0.13 \, \text{J}}{\text{g} \, ^\circ C} \right)(6.2 \, ^\circ C) \]

\[ Z = 6.62 \, \text{J} \]

4) \[ 40.5 \, \text{J} = (15.5 \, \text{kJ})(0.235 \, \text{J}) \left( \frac{0.235 \, \text{J}}{\text{g} \, ^\circ C} \right) \Delta T \]

\[ \Delta T = 11.1 \, ^\circ C \]
\begin{align*}
(5) \quad \lambda &= 25.7 \text{g} + 80.0 \text{g} \\
&= 105.7 \text{g} \\
\varepsilon &= \frac{\text{nc} \Delta T}{2} \\
\varepsilon &= (105.7 \text{g}) \left( \frac{4.1845 \text{ J}}{\text{g} \cdot \text{K}} \right) \left( 24.4 \text{K} - 20.5 \text{K} \right) \\
\varepsilon &= 1724.8 \text{J} \times \frac{1 \text{kJ}}{1000 \text{J}} \\
\Delta H &= -1.72 \text{kJ}
\end{align*}

\begin{align*}
(6) \quad \lambda &= 16.9 \text{g} + 70.0 \text{g} \\
&= 86.9 \text{g} \\
\varepsilon &= \frac{\text{nc} \Delta T}{2} \\
\varepsilon &= (86.9 \text{g}) \left( \frac{4.1845 \text{ J}}{\text{g} \cdot \text{K}} \right) \left( 86.6 \text{K} - 22.4 \text{K} \right) \\
&= 2.334 \times 10^4 \text{J} \times \frac{1 \text{kJ}}{1000 \text{J}} \\
\Delta H &= -23.3 \text{kJ}
\end{align*}
1. From these data,

\[ \text{S(rhombic)} + \text{O}_2(g) \rightarrow \text{SO}_2(g) \quad \Delta H_{\text{rxn}}^\circ = -296.06 \text{ kJ} \]

\[ \text{S(monoclinic)} + \text{O}_2(g) \rightarrow \text{SO}_2(g) \quad \Delta H_{\text{rxn}}^\circ = -296.36 \text{ kJ} \]

calculate the enthalpy change for transformation

\[ \text{S(rhombic)} \rightarrow \text{S(monoclinic)} \]

(Monoclinic and rhombic are different allotropic forms of elemental sulfur.)

2. From the following data,

\[ \text{C(graphite)} + \text{O}_2(g) \rightarrow \text{CO}_2(g) \quad \Delta H_{\text{rxn}}^\circ = -393.5 \text{ kJ} \]

\[ \text{H}_2(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{H}_2\text{O}(l) \quad \Delta H_{\text{rxn}}^\circ = -285.8 \text{ kJ} \]

\[ 2\text{C}_2\text{H}_6(g) + 7\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 6\text{H}_2\text{O}(l) \quad \Delta H_{\text{rxn}}^\circ = -3119.6 \text{ kJ} \]

calculate the enthalpy change for the reaction

\[ 2\text{C(graphite)} + 3\text{H}_2(g) \rightarrow \text{C}_2\text{H}_6(g) \]

3. From the following heats of combustion,

\[ \text{CH}_3\text{OH}(l) + \frac{3}{2}\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(l) \quad \Delta H_{\text{rxn}}^\circ = -726.4 \text{ kJ} \]

\[ \text{C(graphite)} + \text{O}_2(g) \rightarrow \text{CO}_2(g) \quad \Delta H_{\text{rxn}}^\circ = -393.5 \text{ kJ} \]

\[ \text{H}_2(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{H}_2\text{O}(l) \quad \Delta H_{\text{rxn}}^\circ = -285.8 \text{ kJ} \]

calculate the enthalpy of formation of methanol (CH\textsubscript{3}OH) from its elements:

\[ \text{C(graphite)} + 2\text{H}_2(g) + \frac{3}{2}\text{O}_2(g) \rightarrow \text{CH}_3\text{OH}(l) \]

4. Calculate the standard enthalpy change for the reaction

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

given that

\[ 2\text{Al}(s) + \frac{3}{2}\text{O}_2(g) \rightarrow \text{Al}_2\text{O}_3(s) \quad \Delta H_{\text{rxn}}^\circ = -1601 \text{ kJ} \]

\[ 2\text{Fe}(s) + \frac{3}{2}\text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s) \quad \Delta H_{\text{rxn}}^\circ = -821 \text{ kJ} \]
1) \[ S(\text{rhom}) + \frac{3}{2} \text{O}_2 \rightarrow S\text{O}_2 \rightarrow \frac{3}{2} \text{O}_2 + \frac{3}{2} \text{O}_2 \rightarrow 2 \text{S(mero)} + 2 \times 296.36 \] 
\[ S(\text{rhom}) \rightarrow S(\text{mero}) + 0.20 \text{ kJ} \]

2) \[ 2 \text{C} + \frac{1}{2} \text{O}_2 \rightarrow 2 \text{CO}_2 \quad 2 \times (-393.5 \text{ kJ}) \]
\[ 3 \text{H}_2 + \frac{3}{2} \text{O}_2 \rightarrow 3 \text{H}_2\text{O} \quad 3 \times (-285.8) \]
\[ \frac{1}{2} \text{CO}_2 + \frac{5}{2} \text{H}_2\text{O} \rightarrow \frac{7}{2} \text{O}_2 + \frac{5}{2} \text{C}_2\text{H}_6 + 311.9.6 \]
\[ \frac{311.9.6}{2} \]
\[ 2 \text{C} + 3 \text{H}_2 \rightarrow \text{C}_2\text{H}_6 \quad -84.6 \text{ kJ} \]
3) $\text{CO}_2 + 2\text{H}_2\text{O} \rightarrow \frac{3}{2}\text{O}_2 + \text{CH}_3\text{OH}$  

$\text{C} + \text{O}_2 \rightarrow \text{CO}_2$  

$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$  

$\text{C} + 2\text{H}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{CH}_3\text{OH}$

$-258.7 \text{ KJ}$

4) $2\text{Al} + \frac{3}{2}\text{O}_2 \rightarrow \text{Al}_2\text{O}_3$  

$\text{Fe}_2\text{O}_3 \rightarrow 2\text{Fe} + \frac{3}{2}\text{O}_2$  

$2\text{Al} + \text{Fe}_2\text{O}_3 \rightarrow 2\text{Fe} + \text{Al}_2\text{O}_3$  

$-780 \text{ KJ}$
Enthalpy Worksheet

1. The combustion of methane, \( CH_4 \), releases \( 890.4 \text{ kJ/mol} \) of heat. That is, when one mole of methane is burned, \( 890.4 \text{ kJ} \) are given off to the surroundings. This means that the products have \( 890.4 \text{ kJ} \) less energy stored in the bonds than the reactants. Thus, \( \Delta H \) for the reaction = \( -890.4 \text{ kJ} \). A negative symbol for \( \Delta H \) indicates an exothermic reaction.

\[
CH_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(L) \quad \Delta H = -890.4 \text{ kJ}
\]

A. How much energy is given off when 2.00 mol of \( CH_4 \) are burned?

\[
2 \text{ mol} \times \frac{-890.4 \text{ kJ}}{1 \text{ mol}} = -1780 \text{ kJ}
\]

B. How much energy is released when 22.4 g of \( CH_4 \) are burned?

\[
22.4 \text{ g} \times \frac{1 \text{ mol} \text{CH}_4}{16 \text{ g}} \times \frac{-890.4 \text{ kJ}}{1 \text{ mol} \text{CH}_4} = -1246 \text{ kJ}
\]

C. If you were to attempt to make 45.0 g of methane from \( \text{CO}_2 \) and \( \text{H}_2\text{O} \) (with \( \text{O}_2 \) also being produced), how much heat would be absorbed during the reaction?

\[
\text{CO}_2 + 2 \text{H}_2\text{O} \rightarrow \text{CH}_4 + 2 \text{O}_2 \quad \Delta H = +890.4 \text{ kJ}
\]

\[
45.0 \text{ g} \times \frac{1 \text{ mol} \text{CH}_4}{16 \text{ g}} \times \frac{+890 \text{ kJ}}{1 \text{ mol}} = +2503 \text{ kJ}
\]
Use the following heat of formation table in questions 2 – 6.

The Standard Enthalpy and Entropy of Various Substances

<table>
<thead>
<tr>
<th>Substance</th>
<th>$\Delta H_f^\circ$ (kJ/mol)</th>
<th>$S^\circ$ (J/K mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$C_4H_{10(g)}$</td>
<td>$-126$</td>
<td>310</td>
</tr>
<tr>
<td>$CaC_{2(s)}$</td>
<td>$-63$</td>
<td>70</td>
</tr>
<tr>
<td>$Ca(OH)_{2(s)}$</td>
<td>$-987$</td>
<td>88</td>
</tr>
<tr>
<td>$C_2H_2(g)$</td>
<td>227</td>
<td>201</td>
</tr>
<tr>
<td>$CO_2(g)$</td>
<td>$-394$</td>
<td>214</td>
</tr>
<tr>
<td>$H_2(g)$</td>
<td>0</td>
<td>131</td>
</tr>
<tr>
<td>$H_2O(g)$</td>
<td>$-242$</td>
<td>189</td>
</tr>
<tr>
<td>$H_2O(L)$</td>
<td>$-286$</td>
<td>70</td>
</tr>
<tr>
<td>$NH_3(g)$</td>
<td>$-46$</td>
<td>193</td>
</tr>
<tr>
<td>$NO(g)$</td>
<td>90</td>
<td>211</td>
</tr>
<tr>
<td>$NO_2(g)$</td>
<td>34</td>
<td>240</td>
</tr>
<tr>
<td>$N_2O(g)$</td>
<td>82</td>
<td>220</td>
</tr>
<tr>
<td>$O_2(g)$</td>
<td>0</td>
<td>265</td>
</tr>
<tr>
<td>$O_3(g)$</td>
<td>143</td>
<td>239</td>
</tr>
</tbody>
</table>

2. Using data from the heat of formation table above, calculate the enthalpy of reaction for

$$3 \ H_2(g) + O_3(g) \rightarrow 3 \ H_2O(g)$$

$$\Delta H = -869 \ \text{kJ}$$

3. Using data from the heat of formation table above, calculate the heat of reaction for

$$2 \ NO(g) + O_2(g) \rightarrow 2 \ NO_2(g)$$

$$\Delta H = [2 (+84 kJ)] - [2 (90 kJ)]$$

$$\Delta H = -112 \ \text{kJ}$$

4. Using data from the heat of formation table above, calculate the heat of reaction for

$$N_2O(g) + NO_2(g) \rightarrow 3 \ NO(g)$$

$$\Delta H = [3 (90 kJ)] - [82kJ + 34 kJ] + 154 \ \text{kJ}$$
5. Using data from the heat of formation table above, calculate the heat of reaction for
\[ CaC_2(a) + 2 \, H_2O(l) \rightarrow Ca(OH)_2(s) + C_2H_2(g) \]
\[
\left[ -987 \, \text{kJ} + 227 \, \text{kJ} \right] - \left[ -63 \, \text{kJ} + 2 \left( -286 \, \text{kJ} \right) \right] = \Delta H = -125 \, \text{kJ}
\]

6. Many cigarette lighters contain liquid butane, \( C_4H_{10} \). Using the heat of formation table above, calculate the quantity of heat produced when 1.0 g of gaseous butane is completely combusted in air.

\[ 2 \, C_4H_{10} + 13 \, O_2 \rightarrow 10 \, H_2O + 8 \, CO_2 \]
\[
\left[ 8 \left( -394 \, \text{kJ} \right) + 10 \left( -242 \right) \right] - \left[ 2 \left( -126 \, \text{kJ} \right) + 13 \left( 0 \right) \right] = \Delta H = -5320 \, \text{kJ}
\]
Entropy Worksheet

Use the following entropy of formation table in questions 1 – 5.

The Standard Enthalpy and Entropy of Various Substances

<table>
<thead>
<tr>
<th>Substance</th>
<th>$\Delta H^\circ ,(kJ/mol)$</th>
<th>$S^\circ ,(J/K \cdot mol)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$C_4H_{10}(g)$</td>
<td>-126</td>
<td>310</td>
</tr>
<tr>
<td>$CaC_{2}(s)$</td>
<td>-63</td>
<td>70</td>
</tr>
<tr>
<td>$Ca(OH)_{2}(s)$</td>
<td>-937</td>
<td>83</td>
</tr>
<tr>
<td>$C_2H_{2}(s)$</td>
<td>227</td>
<td>201</td>
</tr>
<tr>
<td>$CO_2(g)$</td>
<td>-394</td>
<td>214</td>
</tr>
<tr>
<td>$H_2(g)$</td>
<td>0</td>
<td>131</td>
</tr>
<tr>
<td>$H_2O(g)$</td>
<td>-242</td>
<td>189</td>
</tr>
<tr>
<td>$H_2O(L)$</td>
<td>-286</td>
<td>70</td>
</tr>
<tr>
<td>$NH_3(g)$</td>
<td>-46</td>
<td>193</td>
</tr>
<tr>
<td>$NO(g)$</td>
<td>90</td>
<td>211</td>
</tr>
<tr>
<td>$NO_2(g)$</td>
<td>34</td>
<td>240</td>
</tr>
<tr>
<td>$N_2O(g)$</td>
<td>82</td>
<td>220</td>
</tr>
<tr>
<td>$O_2(g)$</td>
<td>0</td>
<td>205</td>
</tr>
<tr>
<td>$O_3(g)$</td>
<td>143</td>
<td>239</td>
</tr>
</tbody>
</table>

1. Using data from the entropy of formation table above, calculate the entropy of reaction for

$$3\ H_2(g) + O_3(g) \rightarrow 3\ H_2O(g)$$

$$\Delta S^\circ = (3 \cdot 189) - (3 \cdot 131  + 239)$$

2. Using data from the entropy of formation table above, calculate the change in entropy for

$$2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$$

$$\Delta S^\circ = (2 \cdot 240) - (2 \cdot 211 + 205)$$

$$\Delta S^\circ = -147 \ \frac{J}{K}$$
3. Using data from the heat of formation table above, calculate the $\Delta S^\circ$ for

$$N_2O(g) + NO(g) \rightarrow 3 NO(g):$$

$$= 3 \left[ \frac{211}{2} \right] - \left[ \frac{220 + 240}{2} \right]$$

$$\Delta S^\circ = 173 \text{ J/K}$$

4. Using data from the entropy of formation table above, calculate the heat of reaction for

$$CaC_2(s) + 2 H_2O(l) \rightarrow Ca(OH)_2(s) + C_2H_2(g):$$

$$= \left[ \frac{83 + 201}{2} \right] - \left[ \frac{78 + 70}{2} \right]$$

$$\Delta S^\circ = 144 \text{ J/K}$$

5. Using the entropy of formation table above, calculate the change in entropy for the following reaction.

$$C_4H_{10}(g) + \frac{13}{2} O_2(g) \rightarrow 4 CO_2(g) + 5 H_2O(g)$$
1) \[ 3 \text{H}_2 + \text{O}_2 \rightarrow 3 \text{H}_2\text{O} \]

\[ = 3 \left( \frac{189 \text{ J}}{\text{mol} \text{K}} \right) - \left[ 3 \left( \frac{131 \text{ J}}{\text{mol} \text{K}} \right) + 239 \frac{\text{J}}{\text{K}} \right] \]

\[ \Delta S^\circ = -65 \text{ J/K} \]

2) \[ 2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2 \]

\[ = 2 \left( \frac{240 \text{ J}}{\text{mol} \text{K}} \right) - \left[ 2 \left( \frac{211 \text{ J}}{\text{mol} \text{K}} \right) + 205 \frac{\text{J}}{\text{K}} \right] \]

\[ \Delta S^\circ = -147.5 \text{ J/K} \]

3) \[ \text{N}_2\text{O} + \text{NO}_2 \rightarrow 3\text{NO} \]

\[ = 3 \left( \frac{211 \text{ J}}{\text{mol} \text{K}} \right) - \left[ 220 \frac{\text{J}}{\text{K}} + 240 \frac{\text{J}}{\text{K}} \right] \]

\[ \Delta S^\circ = +173 \text{ J/K} \]
4) \[CaC_2 + 2H_2O \rightarrow Ca(OH)_2 + C_2H_2\]

\[= [83.5 + 201.5] - [70.5 + 4(70.5)]\]

\[\Delta S = 74.5 \text{ J/mol K}\]

5) \[C_4H_{10} + \frac{13}{2}O_2 \rightarrow 214 \quad 189\]

\[4CO_2 + 5H_2O\]

\[= [4(214) + 5(189)] - [310 + 15\frac{1}{2}(205)]\]

\[\Delta S = 158.5 \text{ J/mol K}\]
Worksheet on Entropy and Free Energy

1. Define entropy in your own words, and list the variables or conditions that you must consider when comparing the entropy of two substances, or when trying to determine the relative change in entropy.

   production of gas and particles = more disorder at high temps.

2. Does entropy increase or decrease with increase in temperature? Explain.

3. Without reference to any data tables, which member of the following pairs has the greater predicted amount of entropy?

<table>
<thead>
<tr>
<th>A</th>
<th>B</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl (s) or NaCl (aq)</td>
<td>Free CO₂ (g) or CO₂ (g) dissolved in water</td>
</tr>
<tr>
<td>MgS (s) or MgF₂(s)</td>
<td>NO₂ (g) or N₂O₄ (g)</td>
</tr>
<tr>
<td>Ne (g) in a 1 L container or Ne(g) in a 2 L container</td>
<td>Al (s) at 25°C or Al (s) at 100°C</td>
</tr>
</tbody>
</table>

4. For each of the following reactions, indicate whether you would expect the entropy of the system to increase or decrease, and explain why. If you cannot tell just by inspecting the equation, explain why.

   (a) 2 H₂S (g) + 3 O₂ (g) → 2 SO₂ (g) + 2 H₂O (g)  
   \[ \Delta S (-) \]  
   \[ \Delta S (+) \]

   (b) N₂O₄(g) → 2NO₂(g)  
   \[ \Delta S (+) \]  
   \[ \Delta S (-) \]

   (c) CO(g) + H₂O(g) → CO₂(g) + H₂(g)  
   \[ \text{no change in moles of gas} \]  
   \[ \Delta S \]  
   \[ \Delta S (+) \]  
   \[ \Delta S (-) \]

   (d) 2KClO₃(s) → 2KCl(s) + 3O₂(g)  
   \[ \Delta S (+) \]  
   \[ \Delta S (-) \]

   (e) 2NH₃(g) + H₂SO₄(aq) → (NH₄)₂SO₄(aq)  
   \[ \Delta S(-) \]  
   \[ \Delta S(+) \]

5. Consider the reaction below and without reference to any data tables, draw the proper conclusion on which condition A–D best represents the reaction being spontaneous as written.

   \[ \text{CO}_2 (g) + \text{H}_2\text{O} (g) \rightarrow \text{HCOOH} \ (l) \quad \Delta H = -150 \text{ kJ} \quad \Delta S (-) \]

   A. The reaction would be spontaneous at all temperatures.
   B. The reaction would be spontaneous at lower temperatures but not at higher temperatures.
   C. The reaction would be spontaneous at higher temperatures but not at lower temperatures.
   D. The reaction would not be spontaneous at any temperature.

   Briefly explain.

   \[ \Delta G = \Delta H - T \Delta S \]
   \[ (-) \]  
   \[ (-) \]  
   makes \( \Delta G < 0 \) at high temp.
6. Predict which of the following reactions has a positive entropy change.

I. \(2 \text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{N}_2\text{O}(\text{g})\)
II. \(\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})\)
III. \(\text{Zn}(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g})\)

\(\Delta S = +1\)

7. Predict which of the following reactions has a negative entropy change.

I. \(2 \text{HgO}(\text{s}) \rightarrow 2 \text{Hg}(\text{l}) + \text{O}_2(\text{g})\)
II. \(\text{Ba}^2+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})\)
III. \(2\text{H}_2\text{O}_2(\text{l}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})\)

\(\Delta S = -1\)

8. Metallic iron is produced by a high temperature process through a series of reactions, one of which is shown below. Find the various parameters requested using the data which can be found below.

\(3\text{Fe}_2\text{O}_3(\text{s}) + \text{CO}(\text{g}) \rightarrow 2\text{Fe}_3\text{O}_4(\text{s}) + \text{CO}_2(\text{g})\)

What is the standard entropy change for the reaction? Use the table below:

Given Data:

<table>
<thead>
<tr>
<th>Species</th>
<th>(S^\circ) (J/K\cdot mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>\text{CO}_2(\text{g})</td>
<td>213.7</td>
</tr>
<tr>
<td>\text{CO}(\text{g})</td>
<td>197.5</td>
</tr>
<tr>
<td>\text{H}_2\text{O}(\text{g})</td>
<td>188.7</td>
</tr>
<tr>
<td>\text{Fe}_3\text{O}_4(\text{s})</td>
<td>145.3</td>
</tr>
<tr>
<td>\text{H}_2\text{O}(\text{l})</td>
<td>69.94</td>
</tr>
<tr>
<td>\text{Fe}_2\text{O}_3(\text{s})</td>
<td>87.40</td>
</tr>
<tr>
<td>\text{CH}_4\text{O}(\text{l})</td>
<td>127</td>
</tr>
<tr>
<td>\text{C}_5\text{H}_6\text{O}(\text{l})</td>
<td>161</td>
</tr>
<tr>
<td>\text{C}_6\text{H}_6\text{O}(\text{l})</td>
<td>329.3</td>
</tr>
<tr>
<td>\text{C}(\text{diamond})</td>
<td>2.439</td>
</tr>
</tbody>
</table>

\[
\Delta S_{\text{rxn}} = \left[ 213.7 + 2(145.3) \right] - \left[ 3(87.4) + 197.5 \right] = +44.6 \text{ J/K} 
\]
9. Consider the following spontaneous gas phase reaction of $A_2$ molecules (open circles) and $B_2$ molecules (filled circles):

(a) Write a balanced equation for the reaction. Make sure you include states of reactants and products.

(b) What are the signs (+, -, or 0) of $\Delta H$, $\Delta S$, and $\Delta G$ for the reaction? (One of these you should KNOW from given info, one you should easily be able to predict, and the third you predict based on the other two.) Explain your answers.

10. Would you expect each of the following reactions to be spontaneous at low temperatures, high temperatures, all temperatures, or not at all? Explain.

(a) $\text{PCl}_3(g) + \text{Cl}_2(g) \rightarrow \text{PCl}_5(g) \quad \Delta H = -87.9\text{kJ}$

(b) $2\text{NH}_3(g) \rightarrow \text{N}_2(g) + 3\text{H}_2(g) \quad \Delta H = +92.2\text{kJ}$

(c) $2\text{N}_2\text{O}(g) \rightarrow 2\text{N}_2(g) + \text{O}_2(g) \quad \Delta H = -164.1\text{kJ}$

11. Please indicate if TRUE or FALSE:
   a. The entropy of a gas increases with increasing temperature.
   b. Spontaneous processes always increase the entropy of the reacting system.
   c. All spontaneous processes release heat to the surroundings.
   d. An endothermic reaction is more likely to be spontaneous at high temperatures than at low temperatures.
   e. The entropy of sugar decreases as it precipitates from an aqueous solution.

12. Consider the following reaction:
    $\text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)$

    If $\Delta H = -92.2\text{kJ}$ and $\Delta S = -0.1987\text{kJ/K}$, what is $\Delta G$ for the reaction at 475°C and at 5°C? Provide the answer in kJ. In each case, is the reaction spontaneous?

13. The enthalpy and entropy change of a reaction are $-3.9\text{kJ/mole}$ and $+56.6\text{J/mole K}$ respectively at 25°C. What is the free energy change in kJ/mole? Is this reaction always spontaneous, never spontaneous or does it depend on the temperature?
12) \[ \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \]
\[ \Delta G = -92.2 \text{kJ} - (475 + 273k)\left(-\frac{0.1987 \text{kJ}}{k}\right) \]
\[ = +56.4 \text{kJ} \quad \text{non spontaneous} \]
\[ \Delta G = -92.2 \text{kJ} - (5 + 273k)\left(-\frac{0.1987 \text{kJ}}{k}\right) \]
\[ = -37.0 \text{kJ} \quad \text{spontaneous} \]

13) \[ \Delta G \left(\frac{-3.9 \text{kJ}}{\text{mol}}\right) - (25 + 273k) \]
\[ \left(\frac{56.65 \text{kJ}}{\text{mol k} \times \frac{1 \text{kJ}}{1000 \text{J}}}\right) \]
\[ = -20.6 \text{kJ} \quad \text{spontaneous at all temps.} \]