Chapter 17: Section 17.1 Reading Guide

Name ____ **Period** ______

Unit 9 – Worksheet 1 (Goals 1 & 2)

Read page 556 about Energy Transformations

1. Define **Thermochemistry**.

   Study of energy changes that occur in chemical reactions and changes of state.

2. Define "**chemical potential energy**".

   Energy "stored" in chemical bonds (be very careful with this definition – it leads to misunderstanding about energy of chemical bonds). It is better to think of CPE as the energy related to the arrangement of elements inside a chemical compound.

3. During a chemical reaction, a substance is transformed into another substance with a different amount of chemical potential energy. This energy change can occur as either _________**Heat Transfer**_______ or

   ___________**Work**___________ or a _____________**Combination of Both**______

4. Define "**Heat**".

   Energy that transfers between objects due to a temperature difference

5. We defined "**Temperature**" in our last unit. What is the official definition of temperature?

   Average Kinetic Energy of the particles in a substance

6. Describe what happens when two objects of different temperatures come in contact with each other. Explain what is happening in terms of temperature and heat.

   They will exchange heat. The high temperature object will LOSE heat. The low temperature object will GAIN heat. They will exchange heat until they reach the same temperature.

Read page 557 about Endothermic and Exothermic Processes

7. What do we mean by the "system" and the "surroundings" when describing thermochemical processes?

   System: Part of the universe at the focus of our attention (a chemical change or substance experiencing a change)
   Surroundings: The rest of the universe around the system

8. Explain what an endothermic process is. Describe what happens to the "system" and the "surroundings" during an endothermic process.

   Endothermic is when the system is GAINING energy (heat). This heat energy must come from the surroundings.

   The system gains heat, the surroundings lose heat.

9. Explain what an exothermic process is. Describe what happens to the "system" and the "surroundings" during an exothermic process.

   Exothermic is when the system is LOSING energy (heat). This heat energy goes to the surroundings.

   The system loses heat, the surroundings gain heat.

Read Sample Problem 17.1 on page 558

10. On a sunny winter day, the snow on a rooftop begins to melt. As the melted water drips from the roof, it refreezes into icicles. Describe the direction of heat flow as the water freezes. Is this process endothermic or exothermic?

    Heat flows out of the water as it freezes (into the surroundings). This is EXOTHERMIC

The following are problems 1 and 2 found on page 558. You can check your work in the appendix at the back of the textbook.

1. *A container of melted wax stands at room temperature. What is the direction of heat flow as the liquid wax solidifies? Is the process endothermic or exothermic?*

   Heat flows to the surroundings. EXOTHERMIC

2. *When barium hydroxide octahydrate [$Ba(OH)_2 \cdot 8H_2O$] is mixed in a beaker with ammonium thiocyanate [NH_4SCN], a reaction occurs. The beaker becomes very cold. Is the reaction endothermic or exothermic?*

   Heat flows out of the surroundings. ENDOTHERMIC

Read page 558 about Units for Measuring Heat Flow.

11. How do calories, dietary Calories, and Joules relate to each other?

    Calories are the energy to heat 1 gram of water 1°C. 1000 calories are 1 food calorie (kilocalorie). 4.18 Joules is 1 calorie

12. Define each process as endothermic or exothermic.
    a. condensing steam (water/steam is the system): **Exothermic**
    b. evaporating alcohol (alcohol is the system): **Endothermic**
    c. burning alcohol (alcohol and oxygen reactants are system) **Exothermic**
    d. baking a potato (potato is the system) **Endothermic**

Heat Capacity

Name ____ **Period** ______

Unit 9 – Worksheet 2 (Goals 3 & 4)

Answer each question completely.

1. Define energy and explain how energy and heat are related.

   Energy – capacity to do work or supply heat.

   Heat is a form of energy transfer from hot (high energy) to cold (lower energy).

2. Will the specific heat of 50 g of a substances be the same as, or greater than, the specific heat of 10 g of the same substance?

   The specific heat would be the same since specific heat is defined relative to 1 gram of the substance (The heat capacity of the 50 g sample would be greater).

3. On a sunny day, why does the concrete deck around an outdoor swimming pool become hot, while the water stays cool?

   The water has a higher specific heat capacity than concrete. Water takes more energy to cause a change in temperature (average kinetic energy)

4. How much heat in kilojoules is released when 25.0 g of water is cooled from 85.0°C to 40.0°C? The specific heat of water is 4.18 J/g°C.

   $$\boldsymbol{q = mc\ \ T = 25.0\ \mathrm{g}\ \ 4.18\ \frac{\mathrm{J}}{\mathrm{g}\ \ ^{o}\mathrm{C}}\ \ (40.0\ \ \ \ 85.0)\ ^{o}\mathrm{C} =\ \ \ 4.70\mathrm{x}10^{3}\ \mathrm{J\ or}\ \ \ \ 4.70\ \mathrm{kJ}}$$

5. A metal weighing 50.0 g absorbs 220.0 J of heat when its temperature increases by 120.0°C. What is the specific heat of the metal?

   $$\boldsymbol{c = \frac{q}{m\ \ T} = \frac{220.0\ J}{50.0\ g\ \ \ 120.0\ ^{o}\mathrm{C}} = 0.0367\ \frac{\mathrm{J}}{\mathrm{g}\ \ \ ^{o}\mathrm{C}}}$$

6. Calculate the heat gained by 125.0 g of water when it is put into a calorimeter and its temperature is increased by 90.0°C. The specific heat of water is 4.18 J/g°C.

$$q = mc\ \ T = 125.0\ g\ \ \ 4.18\ \frac{\mathrm{J}}{\mathrm{g}\ \ {}^{o}\mathrm{C}}\ \ \ 90.0\ {}^{o}\mathrm{C} = 4.70\mathrm{x}10^{4}\ \mathrm{J\ or}\ 4.70\ \mathrm{kJ}$$

7. When a 50.0-g nugget of pure gold is heated from 35.0°C to 50.0°C, it absorbed 96.8 J of energy. Find the specific heat of gold.

$$c = \frac{q}{m\ \ T} = \frac{96.8\,J}{50.0\,g\ \ \ (50.0\ \ \ \ 35.0)\ {}^{o}\mathrm{C}} = 0.129\ \frac{\mathrm{J}}{\mathrm{g}\ \ {}^{o}\mathrm{C}}$$

8. A 20.0 g sample of aluminum is cooled 7.5⍰C. The specific heat capacity of aluminum is 0.900 J/g⍰C. What is the energy change for this sample?

$$q = mc\ \ T = 20.0\ g\ \ \ 0.900\ \frac{\mathrm{J}}{\mathrm{g}\ \ {}^{o}\mathrm{C}}\ \ \ \ 7.5\ {}^{o}\mathrm{C} = \ \ \ 140\ \mathrm{J}$$

9. 150.0 mL of water absorbs 7.84 kJ of energy. The specific heat capacity of water is 4.18 J/g⍰C. What is the temperature change?

$$T = \frac{q}{mc} = \frac{7.84\ kJ\left(\frac{1000\,J}{1\ kJ}\right)}{150.0\ ml\left(\frac{1\ g}{1\ ml}\right)\ \ 4.18\frac{\mathrm{J}}{\mathrm{g}\ \ {}^{o}\mathrm{C}}} = 12.5\ \ {}^{o}\mathrm{C}$$

10. The specific heat capacity of diamond is 0.5050 J/g⍰C. How much energy is required to heat 25.0 g of diamond from 10.5⍰C to 15.6⍰C?

$$q = mc\ \ T = 25.0\ g\ \ \ 0.5050\ \frac{\mathrm{J}}{\mathrm{g}\ \ {}^{o}\mathrm{C}}\ \ (15.6\ \ \ \ 10.5)\ {}^{o}\mathrm{C} = 64\ \mathrm{J}$$

11. A 26.6 g sample of mercury is heated to 110.0°C and then placed in 125 g of water in a coffee-cup calorimeter. The initial temperature of the water is 23.0°C. The specific heat capacity of water is 4.18 J/g°C, and the specific heat capacity of mercury is 0.139 J/g°C. What is the final temperature of the water and the mercury?

$$q_{Hg} = q_{H_2O}$$

$$mc(T_f\ \ \ \ T_i)_{Hg} = mc(T_f\ \ \ \ T_i)_{H_2O}$$

$$26.6\ g\ \ 0.139\frac{\mathrm{J}}{\mathrm{g}\ \ {}^{o}\mathrm{C}}(T_f\ \ \ \ 110.0\ \ {}^{o}\mathrm{C})_{Hg} = 125\ g\ \ \ 4.18\frac{\mathrm{J}}{\mathrm{g}\ \ {}^{o}\mathrm{C}}(T_f\ \ \ \ 23.0\ \ {}^{o}\mathrm{C})_{H_2O}$$

$$T_f = 23.6\ \ {}^{o}\mathrm{C}$$

Thermochemical Equations and Stoichiometry

Name ____ **Period** ______

Unit 9 – Worksheet 3 (Goal 5)

1. Classify each of these reactions as exothermic or endothermic:

   a. energy + $SO_2(g)$ → S (g) + $O_2(g)$ **Endothermic**

   b. $PCl_3(l)$ + $Cl_2(g)$ → $PCl_5(g)$ + energy **Exothermic**

   c. $NH_3(g)$ + HCl(g) → $NH_4Cl(g)$ + energy **Exothermic**

   d. 2 $KNO_3(s)$ + energy → 2 $KNO_2(s)$ + $O_2(g)$ **Endothermic**

   e. 2 $Al_2O_3(s)$ → 4 Al(s) + 3 $O_2(g)$ H = +3351.4 kJ **Endothermic**

   f. $SnCl_2(s)$ + $Cl_2(g)$ → $SnCl_4(s)$ H = -186.2 kJ **Exothermic**

   g. C(s) + $H_2O(g)$ → CO(g) + $H_2(g)$ H = +131.3 kJ **Endothermic**

   h. 2 $H_2(g)$ + $O_2(g)$ → 2 $H_2O(g)$ H = - 572 kJ **Exothermic**

2. On average, each person in the USA consumes 110 g of sugar daily. If all of this sugar is assumed to be glucose, how much heat energy from glucose is transferred to one person per day?

$$C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O + 2870\ kJ$$

$$\mathbf{110\ g\ C_6H_{12}O_6\left(\frac{1\ mol\ C_6H_{12}O_6}{180.0\ g\ C_6H_{12}O_6}\right)\left(\frac{2870\ kJ}{1\ mol\ C_6H_{12}O_6}\right) = 1800\ kJ}$$

3. A pop contains 41 g of sugar (glucose) in a 12 oz can. How much energy (kJ) in the form of exercise does a person have to do to use the energy up?

$$\mathbf{41\ g\ C_6H_{12}O_6\left(\frac{1\ mol\ C_6H_{12}O_6}{180.0\ g\ C_6H_{12}O_6}\right)\left(\frac{2870\ kJ}{1\ mol\ C_6H_{12}O_6}\right) = 650\ kJ}$$

4. When suffering from a fever, your body temperature rises from 37°C to 40°C, using 787 kJ of energy in the process. Assume that your body burns only glucose to raise your temperature. How many grams of glucose are consumed?

$$\mathbf{787\ kJ\left(\frac{1\ mol\ C_6H_{12}O_6}{2870\ kJ}\right)\left(\frac{180.0\ g\ C_6H_{12}O_6}{1\ mol\ C_6H_{12}O_6}\right) = 49.4\ g\ C_6H_{12}O_6}$$

5. What is the minimum mass of gasoline (C_8H_{18}) that must be delivered by the carburetor or fuel injection system to a car's engine in order to generate 5000 kJ of energy needed to pass a car?

$$2\ C_8H_{18}\ +\ 25\ O_2\ \rightarrow\ 16\ CO_2\ +\ 18\ H_2O\ +\ 10{,}990\ kJ$$

$$\mathbf{5000\ kJ\left(\frac{2\ mol\ C_8H_{18}}{10990\ kJ}\right)\left(\frac{114.0\ g\ C_8H_{18}}{1\ mol\ C_8H_{18}}\right) = 100\ g\ C_8H_{18}}$$

6. A butane (C_4H_{10}) lighter needs 77 kJ to heat some water. What mass of butane is needed?

$$2\ C_4H_{10}\ +\ 13\ O_2\ \rightarrow\ 8\ CO_2\ +\ 10\ H_2O\ +\ 5713\ kJ$$

$$\mathbf{77\ kJ\ \left(\frac{2\ mol\ C_4H_{10}}{5713\ kJ}\right)\left(\frac{58\ g\ C_4H_{10}}{1\ mol\ C_4H_{10}}\right) = 1.6\ g\ C_4H_{10}}$$

7. When 2 mol of solid magnesium combines with one mol of gaseous oxygen, 2 mol of magnesium oxide is formed and 1204 kJ of energy is released. Write the thermochemical equation for this combustion reaction.

2 Mg(s) + O_2(g) → 2 MgO(s) + 1204 kJ or H = -1204 kJ

Heat and Changes of State

Name ____ **Period** ______

Unit 9 – Worksheet 4 (Goal 6)

Use the following table to answer the questions below:

	m.p. °C	b.p. °C	c solid J/g°C	c liquid J/g°C	c gas J/g°C	H_{fus} J/g	H_{vap} J/g
water	0.0	100.0	2.06	4.18	2.02	333	2260
ethanol	-117	78		2.46	0.954	109	855
cesium	28	678	0.246	0.252	0.156	15.7	514
benzene	5.5	80.0		1.74	1.04	127	395
oxygen	-218	-183			0.916	445	6810
gold	1064	2807	0.128	0.150		64	
hydrogen	-259	-253		3.4	14.3		117
copper	1083	2570	0.385			205	4799

1. How much heat energy is required to melt 25 g of Cu?

$$\boldsymbol{m \;\; H_{fus} = 25\,g\left(\frac{205\,J}{g}\right) = 5100\,J}$$

2. How much heat energy is required to melt 10.0 g of Cu metal if the Cu is at room temperature (25.0°C)?

$$\boldsymbol{mc_s\Delta T + m\Delta H_{fus} = 10.0\,g\left(\frac{0.385\,J}{g\,{}^oC}\right)(1083 \quad 25.0)\,{}^oC + 10.0\,g\left(\frac{205\,J}{g}\right) = 6120\,J}$$

3. How much heat energy is required to heat 5.00 g of hydrogen from 25.0°C to 125.0°C?

$$\boldsymbol{mc_g\Delta T = 5.00\,g\left(\frac{14.3\,J}{g\,{}^oC}\right)(125.0 \quad 25.0)\,{}^oC = 7150\,J}$$

4. How much heat energy must be added to 2.00 g of gaseous hydrogen to take it from its boiling point to room temperature (25.0°C)?

$$\boldsymbol{mc_g\Delta T = 2.00\,g\left(\frac{14.3\,J}{g\,{}^oC}\right)(25.0 \qquad 253)\,{}^oC = 7950\,J}$$

5. How much heat energy is required to heat 45.3 g of cesium from 15.0°C to 453.0°C?

$$mc_s\Delta T + m\Delta H_{fus} + mc_l\Delta T =$$
$$45.3\ g\left(\frac{0.246\ J}{g\ ^oC}\right)(28.0 \quad 15.0)\ ^oC + 45.3\ g\left(\frac{15.7\ J}{g}\right) + 45.3\ g\left(\frac{0.252\ J}{g\ ^oC}\right)(453.0 \quad 28.0)\ ^oC = 5{,}710\ J$$

6. How much heat energy is required to convert 20.0 g of ice at a temperature of -15°C to steam at a temperature of 110°C?

$$mc_s\Delta T + m\Delta H_{fus} + mc_l\Delta T + m\Delta H_{vap} + mc_g\Delta T =$$
$$20.0\ g\left(\frac{2.06\ J}{g\ ^oC}\right)(0.0 \quad 15)\ ^oC + 20.0\ g\left(\frac{333\ J}{g}\right) + 20.0\ g\left(\frac{4.18J}{g\ ^oC}\right)(100.0 \quad 0.0)\ ^oC + 20.0\ g\left(\frac{2260\ J}{g}\right)$$
$$+ 20.0\ g\left(\frac{2.02J}{g\ ^oC}\right)(110 \quad 100.0)\ ^oC = 61{,}200\ J$$

7. How much heat energy is used to warm 15 g of ethanol from 10.0°C to 60.0°C?

$$mc_l\Delta T = 15\ g\left(\frac{2.46\ J}{g\ ^oC}\right)(60.0 \quad 10.0)\ ^oC = 1{,}800\ J$$

8. How much heat energy is released when 42 g of steam at 115.0°C is cooled to ice at 0.0°C?

$$mc_g\Delta T + m\Delta H_{con} + mc_l\Delta T + m\Delta H_{sol} =$$
$$42\ g\left(\frac{2.02\ J}{g\ ^oC}\right)(100.0 \quad 115.0)\ ^oC + 42\ g\left(-\frac{2260\ J}{g}\right) + 42\ g\left(\frac{4.18J}{g\ ^oC}\right)(0.0 \quad 100.0)\ ^oC + 42\ g\left(-\frac{333\ J}{g}\right)$$
$$= \quad 128{,}000\ J$$

9. How much heat energy is required to heat 25 g of benzene from 6.0°C to 80.0°C?

$$mc_l\Delta T = 25\ g\left(\frac{1.74\ J}{g\ ^oC}\right)(80.0 \quad 6.0)\ ^oC = 3{,}200\ J$$

10. How much heat energy is required to melt 84 g of oxygen at -218.0°C?

$$m\Delta H_{fus} = 84\ g\left(\frac{445\ J}{g}\right) = 37{,}000\ J$$

Hess's Law

Name ____ **Period** ______

Unit 9 – Worksheet 5 (Goal 7)

1. What is the enthalpy change for the following reaction?

   $Sn(s) + 2\ Cl_2(g) \rightarrow SnCl_4(l)$

   Use the following:

$Sn(s) + Cl_2(g) \rightarrow SnCl_2(s)$	H = -325 kJ
$SnCl_2(s) + Cl_2(g) \rightarrow SnCl_4(l)$	H = -186 kJ

   $1) + 2) = \quad 325\ kJ + \quad 186\ kJ = \quad \mathbf{511\ kJ}$

2. What is the enthalpy change for the following reaction?

   $Mn(s) + O_2(g) \rightarrow MnO_2(s)$

   Use the following:

$MnO_2(s) + Mn(s) \rightarrow 2\ MnO(s)$	H = -240 kJ
$2\ MnO_2(s) \rightarrow 2\ MnO(s) + O_2(g)$	H = +264 kJ

   $1) \quad 2) = \quad 240\ kJ \quad 264\ kJ = \quad \mathbf{504\ kJ}$

3. What is the enthalpy change for the following reaction?

   $SnO_2(s) + 2\ H_2(g) \rightarrow Sn(s) + 2\ H_2O(l)$

   Use the following:

$Sn(s) + O_2(g) \rightarrow SnO_2(s)$	H = -580.7 kJ
$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$	H = -571.6 kJ

   $1) + 2) = \qquad 580.7\ kJ + \quad 571.6\ kJ = \mathbf{9.1\ kJ}$

4. What is the enthalpy change for the following reaction?

   $2\ Mg(s) + SiCl_4(l) \rightarrow Si(s) + 2\ MgCl_2(s)$

   Use the following:

$SiCl_4(l) \rightarrow Si(s) + 2\ Cl_2(g)$	H = +687 kJ
$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$	H = -641 kJ

   $1) + 2 \quad 2) = 687\ kJ + 2 \quad 641\ kJ = \quad \mathbf{595\ kJ}$

5. What is the enthalpy change for the following reaction?

$N_2(g) + 2\,H_2(g) \rightarrow N_2H_4(l)$

Use the following:

$H_2(g) + 1/2\,O_2(g) \rightarrow H_2O(g)$	H = -242 kJ
$N_2(g) + 2\,H_2O(g) \rightarrow N_2H_4(l) + O_2(g)$	H = +534 kJ

$$2\quad 1) + 2) = 2\qquad 242\ kJ + 534\ kJ = \mathbf{5.0\ x\ 10^1\ kJ}$$

6. What is the enthalpy change for the following reaction?

$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(l)$

Use the following:

$C(s) + 2\,H_2(g) \rightarrow CH_4(g)$	H = -74.8 kJ
$C(s) + O_2(g) \rightarrow CO_2(g)$	H = -393.5 kJ
$H_2(g) + 1/2\,O_2(g) \rightarrow H_2O(l)$	H = -235.8 kJ

$$1) + 2) + 2\quad 3) = \qquad 74.8\ kJ + \quad 393.5\ kJ + 2\qquad 235.8\ kJ = \quad \mathbf{790.3\ kJ}$$

7. Calculate the enthalpy of change for:

$2\,C_2H_6(g) + 7\,O_2(g) \rightarrow 4\,CO_2(g) + 6\,H_2O(g)$

Use the following:

$2\,C(s) + 3\,H_2(g) \rightarrow C_2H_6(g)$	H = -84.7 kJ
$C(s) + O_2(g) \rightarrow CO_2(g)$	H = -393.5 kJ
$H_2(g) + 1/2\,O_2(g) \rightarrow H_2O(g)$	H = -241.8 kJ

$$2\quad 1) + 4\quad 2) + 6\quad 3) = \quad 2\qquad 84.7\ kJ + 4\qquad 393.5\ kJ + 6\qquad 241.8\ kJ = \quad \mathbf{2855.4\ kJ}$$