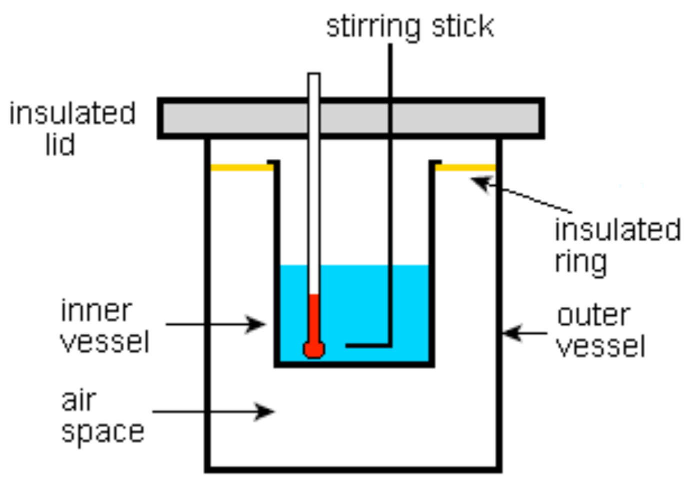
Chemistry B Thermochemistry Packet



Thermochemistry Learning Goals

- **Worksheet 3** (Concept) I can classify chemical reactions as either exothermic (release energy to the environment) or endothermic (absorb energy from the environment).

- **Worksheet 3** (Concept) I can draw enthalpy diagrams (graphs) for exothermic and endothermic reactions.

- **Worksheet 3** (Math) I can use stoichiometry to calculate the amount of heat produced for a given mass of reactant from a balanced chemical equation.

- **Worksheet 4** (Math) I can calculate the ΔH for a given reaction using Hess's Law.

- **Worksheet 5** (Concept) I can describe the relationship between specific heat of a material and the amount of time it take to increase the heat of the material.

- **Worksheet 5** (Math) I can calculate specific heat, temperature change or mass of a material when provided with the appropriate equations.

- **Worksheet 6** (Math) I can calculate the ΔH for a chemical reaction using simple coffee cup calorimetry.

- **Worksheet 7** (Math) I can calculate specific heat AND calorimetry when given a mixture of problems.

- **Worksheet 8** (Concept) I can describe the two natural driving forces: (1) toward minimum energy (enthalpy) and (2) toward maximum disorder (entropy).

Worksheet 1: Types of Energy

Thermochemistry is the study of heat (energy) changes during chemical and physical reactions. As chemists we define **energy** as anything that **can cause a change in the position or properties of matter**. We classify all forms of energy into two general categories – potential energy (stored energy) and kinetic energy (energy due to motion). Within each category, there are many different forms of energy.

FORMS OF ENERGY

POTENTIAL ENERGY Stored energy and energy of position	**KINETIC ENERGY** The motion of waves, electrons, atoms, molecules and substances
↓	↓
CHEMICAL ENERGY Stored in the bonds of atoms and molecules. Petroleum, natural gas, propane, and coal are examples.	**RADIANT ENERGY (LIGHT)** Electromagnetic energy that travels in waves. Solar energy is an example.
NUCLEAR ENERGY Stored in the nucleus of an atom – the energy that holds the nucleus together. The energy in the nucleus of a uranium atom is an example.	**THERMAL ENERGY (HEAT)** Internal energy in substances – the vibration or movement of atoms and molecules in substances. Geothermal is an example.
STORED MECHANICAL ENERGY Stored in objects by the application of force. Compressed springs and stretched rubber bands are examples.	**MOTION** Movement of a substance from one place to another. Wind and hydropower are examples.
GRAVITATIONAL ENERGY Energy of place or position. Water in a reservoir behind a hydropower dam is an example.	**SOUND** Movement of energy through substances in waves.
	ELECTRICAL ENERGY Movement of electrons. Lightning and electricity are examples.

The law of conservation of energy states that energy does not disappear, it just changes forms. This means that energy is neither created nor destroyed.

There is an important relationship between the terms energy, heat, temperature and kinetic energy. In order to understand this relationship, consider the following scenario. There are two containers of sand, one filled about one-thirds of the way and the other filled to the top. Since they have been in the same environment, they should be the same temperature. Each student in class shakes both containers and the final temperatures are recorded. **Predict how the temperatures will compare: ____**

Shaking the container, we moved the grains of sand, causing them to collide with each other. They rubbed against each other, causing friction between the particles and producing thermal energy (heat). We transformed motion energy into thermal energy (heat).

The increase in temperature was greater for the container that was only one-third full, because the partially full container has more space in it for the sand to move around. The grains of sand collide with each other with greater speed. More speed means more kinetic energy (motion). In the container filled with sand, the grains of sand have little space to move. **Considering the previous example, explain in a few sentences how energy, heat, temperature and kinetic energy are related: ____**

____

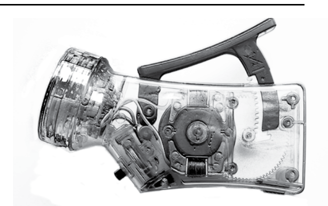
____

Whether we realize it or not, we experience energy transformations- conversions from one type of energy to another, on a regular basis. Consider the transformations in a cell phone. When the phone is charging electrical energy is transformed to chemical energy in order to be stored in a battery. This chemical energy can be transformed to light energy (screen), sound energy (ring tone), motion (vibration).

In the original example the energy flow from shaking the containers of sand originated from our sun. Let's start with the thermal energy, or heat, we produced as a result of the motion our bodies provided. Our bodies got the energy to shake the container from the food we ate. The energy was stored in the food as chemical energy in the bonds of the molecules. Chemical energy is stored in food through photosynthesis. Sunlight, or radiant energy, changes water and carbon dioxide into the glucose and oxygen in plants. Glucose is a sugar with lots of energy—chemical energy.
Nuclear energy produced the radiant energy from the sun. Inside the sun's core, atoms of hydrogen are fused together to form atoms of helium. The fusion occurs because of the tremendous heat and gravitational forces inside the sun's core. The resulting atom of helium has less mass than the original atoms of hydrogen. This missing mass is changed into energy—radiant energy.
Accordingly, the increase in the temperature of the sand is the result of nuclear energy. In fact all energy transformations can be traced back to nuclear energy: fission, fusion, or radioactive decay. The energy stored in fossils fuels (coal, petroleum, natural gas, propane) is a result of sunlight from millions of years ago. Wind, hydropower, and biomass energy are also a result of the sun's radiant energy. Geothermal energy is a result of the radioactive decay of elements in the Earth's core. The electricity produced in a nuclear power plant is the result of the splitting, or fission, of heavy uranium atoms into lighter atoms. When fission occurs, mass is changed into energy. All energy transformations can be traced back to nuclear energy.

Think of a "real-life" example of the following energy transformations:

- Mechanical → Motion ____
- Electrical → Sound ____
- Radiant (Light) → Electrical ____
- Chemical → Thermal (Heat) ____
- Gravitational →Motion ____
- Nuclear → Thermal (Heat) ____
- Chemical → Motion ____
- Sound → Radiant (Light) ____



Now let's imagine another scenario. Some of you may be familiar with a hand generated flashlight. This type of flashlight does not require batteries. Instead, the user shakes the flashlight or pumps a handle in order to generate light. Review the various forms of energy described on the previous page. **List all of energy transformations that must occur in order to generate light:**

____

____

A hand generated flashlight works by converting motion into electrical energy. Electricity is powering the light. The handle is connected to a gear that spins a metal disk, which is a magnet. Above and below the disk are two thin coils of copper wire with two wires leading to a capacitor and LED light bulbs. The magnet spins inside the coils of wire, creating an electric current. This electricity is stored in the capacitor as electrical energy, then is released to power the light bulb. As the handle is squeezed, an electric current is generated that flows to the capacitor, which begins to store the electrical energy. The light will stay on as long as the capacitor is charged. When the handle is squeezed quickly, the light is much brighter. The more energy put into the system, the more energy is stored in the capacitor and the more energy there is to produce light.

Use the flashcards provided to map out all of the energy transformations described above in the correct order.

Worksheet 2: Sources of Energy

All of us use many forms of energy in our daily lives. In order to be well informed citizens, it is important to understand where the energy that we use comes from and what impact our energy use has on the world around us. Recently, many people have become concerned with “energy conservation” in order to avoid an “energy crisis”. In this worksheet we will explore these ideas and what they really mean. To begin, consider the following questions:

1. If energy cannot be destroyed why do we worry about “saving” or “conserving” it?

2. If energy doesn’t disappear why do people worry about an energy crisis?

Let’s begin by distinguishing between nonrenewable and renewable resources. **Nonrenewable resources either cannot be replaced or would take millions of years to replace.** Nonrenewable energy sources are petroleum, natural gas, coal, propane, and uranium. **Renewable resources can be easily replaced over a short period of time.** Renewable energy resources are wind, geothermal, solar, hydropower, and biomass.

3. Using the **Forms of Energy** listed on Worksheet 1, and the graphic below, name the form of energy for each of the sources of energy listed. Remember, if the source of energy must be burned, the form of energy is chemical energy.

RENEWABLE	**NONRENEWABLE**
Biomass ________________________	Petroleum ________________________
Hydropower ________________________	Coal ________________________
Wind ________________________	Natural Gas ________________________
Geothermal ________________________	Uranium ________________________
Solar ________________________	Propane ________________________

4. Look at the U.S. Energy Consumption by Source graphic below and calculate the percentage of the nation’s energy use that each form of energy provides.

What percentage of the nation’s energy is provided by each form of energy?

Chemical _____

Nuclear _____

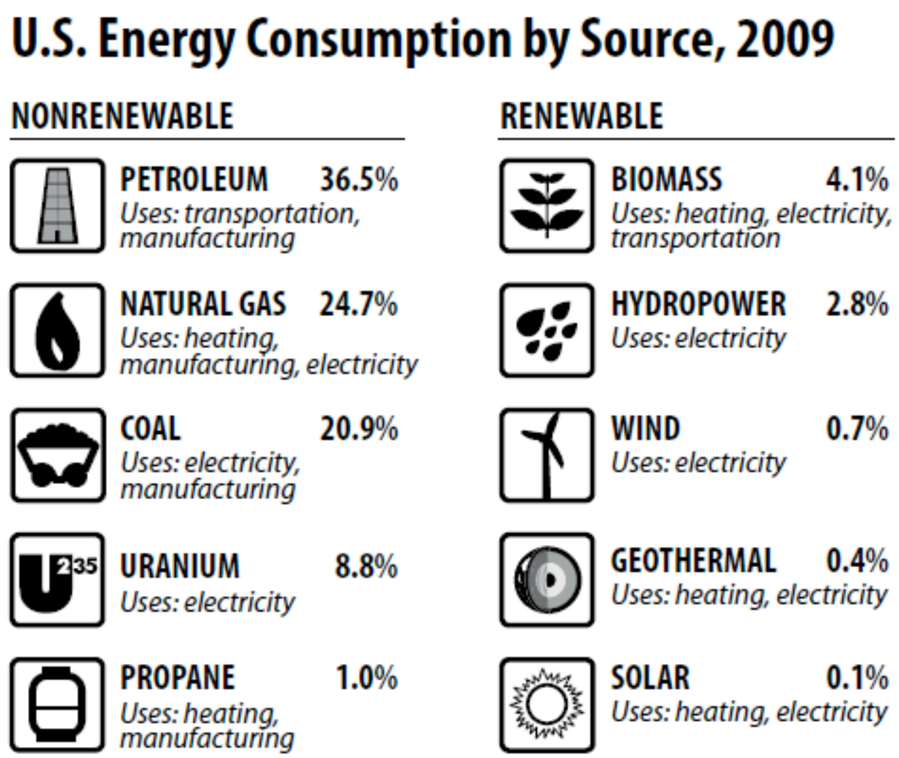
Motion_____

Thermal _____

Radiant _____

What percentage of the nation’s energy is provided by renewables? ______

By nonrenewables? ______



Source: http://www.need.org/needpdf/Science%20of%20Energy.pdf

Worksheet 3: Enthalpy

Almost all chemical and physical reactions involve energy (usually in the form of heat) being released or added. **An exothermic change is a reaction that releases energy. An endothermic change is one in which the energy must be added for the reaction to occur.** For **exothermic** reactions, **energy can be thought of as a product** in the reaction.For **endothermic** changes, **energy can be thought of as a reactant** in the reaction.

If a chemical reaction occurs at constant pressure, as all of our chemical reactions do we can consider a property called enthalpy. **Enthalpy (H) is the energy (heat) content of a system at constant pressure.** You cannot measure the actual energy or enthalpy of a substance, but you can measure the **change in enthalpy** for a reaction. This change is symbolized by **ΔH.**

For exothermic reactions, enthalpy values are always negative, that is the energy of the products is lower than that of the reactants. This is because energy is released as new bonds are formed in the products and this amount of energy is greater than the energy required to break the old bonds in the reactants.

$\Delta H = H_{products} - H_{reactants}$ ($_{small\ \#}$ - BIG#) = - negative #

For example: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$ **ΔH = -1625 kJ**

For endothermic reactions, enthalpy values are always positive, that is that energy of the products is greater than that of the reactions. This is because the energy released as new bonds are formed in the products is less than the energy required to break the bonds in the reactants. This energy must be supplied in order for the reaction to occur. The added energy does not disappear, of course due to the Law of Conservation of Energy. Instead, it becomes stored in the chemical bonds of the products.

$\Delta H_{rxn} = H_{products} - H_{reactants}$ (BIG# - $_{small\ \#}$) = positive #

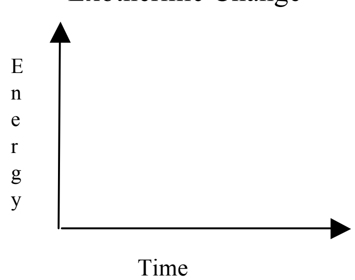
For example: $C + H_2O \rightarrow CO + H_2$ **ΔH = + 113 kJ**

Complete the following chart:

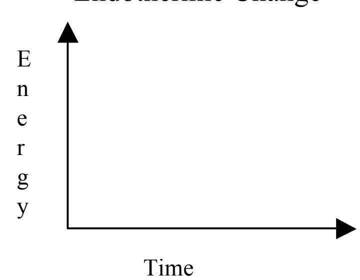
Type of Reaction	**Sign of ΔH**	**Which has more energy: reactants or products?**
Exothermic		
Endothermic		

Make enthalpy diagrams for the two chemical reactions from above, showing clearly the amount of energy released or gained.

Exothermic Change



Endothermic Change



Does the energy go from the surrounding to the chemicals or from the chemicals to the surroundings in an exothermic reaction?

Classify each of the following as an exothermic or endothermic process.

Melting ice cubes ______________________

Burning a candle _____________________

Evaporation of water_____________________

Baking Bread ______________________

Splitting a gas molecule apart __________________

Formation of snow in clouds ____________________

Chemistry problems involving enthalpy changes are similar to the stoichiometry problems that we learned how to do earlier in the trimester. The amount of energy that is absorbed or released in a reaction depends on the number of moles of reactants involved.

For example, according to the following equation, 113 kJ of heat is absorbed when **one mole** of carbon reacts with **one mole** of water.

$C + H_2O \rightarrow CO + H_2 \quad \Delta H = +113 \text{ kJ}$

How much heat will be absorbed if 2 moles of carbon combine with 2 moles of water? If you said ________kJ of heat are absorbed—twice the amount absorbed by one mole of each of the reactants you are correct! The enthalpy change for a reaction is proportionately smaller or larger depending on the amounts of the reactants and products involved.

Example Problem:

How much heat will be released if 1.0 gram of hydrogen peroxide H_2O_2 decomposes.

$2H_2O_2 \rightarrow 2H_2O + O_2 \quad \Delta H = -190 \text{ kJ}$

$\#\text{kJ} = 1.0 \text{ grams } H_2O_2 \times \frac{1 \text{ mol } H_2O_2}{34.02 \text{ g } H_2O_2} \times \frac{-190 \text{ kJ}}{2 \text{ moles}} = -2.7924 \text{ kJ} = -2.8 \text{ kJ}$ (the minus sign tells you heat is released)

Solve the following problems on a separate sheet of paper showing all work.

1. How much heat is released when 9.22 grams of glucose $C_6H_{12}O_6$ in your body reacts with according to the following equation?

$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O \quad \Delta H = -2803 \text{ kJ}$

2. How much heat is absorbed during photosynthesis when 9.22 grams of glucose $C_6H_{12}O_6$ is produced?

$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2 \quad \Delta H = +2803 \text{ kJ}$

3. How much heat is released when 147 grams of NO_2 is dissolved in excess water?

$3NO_2 + H_2O \rightarrow 2HNO_3 + NO \quad \Delta H = -138 \text{ kJ}$

4. Calculate the heat released when 74.6 grams of SO_2 reacts according to the following equation.

$2SO_2 + O_2 \rightarrow 2SO_3 \quad \Delta H = -99.1 \text{ kJ}$

5. Calculate the heat release when 266 grams of white phosphorus P_4 burns in air according to the following equation.

$P_4 + 5O_2 \rightarrow P_4O_{10} \quad \Delta H = -3013 \text{ kJ}$

6. Calculate the amount of heat released when 1.26×10^4 grams of ammonia NH_3 are produced according to the following equation.

$N_2 + 3H_2 \rightarrow 2NH_3 \quad \Delta H = -92.6 \text{ kJ}$

Worksheet 4: Hess's Law

Suppose you are studying the formation of acid rain that results from the reaction of water in the atmosphere with sulfur trioxide given off from a volcanic eruption. You would need to determine the enthalpy (ΔH) for the following reaction:

$2S + 3O_2 \rightarrow 2SO_3$ $\Delta H = ?$

Unfortunately, this reaction cannot be duplicated in lab in the same way it occurs in nature.

In situations like this, you can calculate the enthalpy using Hess's Law of Heat Summation. **Hess's Law says that if a series of reactions are added together, the enthalpy change for the net reaction will be the sum of the enthalpy changes for the individual steps.** This law enables you to calculate enthalpy changes for an enormous number of chemical reactions by imagining that each reaction occurs through a series of steps for which the enthalpy changes are know. I**n these types of problems, you let the overall or net equation guide you in your manipulations of the step equations.**

Example Problem: Given the following chemical equations, calculate the energy change for the reaction that produces SO_3.

(A) $S + O_2 \rightarrow SO_2$ $\Delta H = -297\ kJ$

(B) $2SO_3 \rightarrow 2SO_2 + O_2$ $\Delta H = 198\ kJ$

The **desired chemical equation is: $2S + 3O_2 \rightarrow 2SO_3$**

Since the desired equation shows two moles of sulfur reacting, equation (A) must be doubled, multiplying all of the coefficients by 2. **When you double a reaction, ΔH must also be doubled** because twice the energy will be released. Applying these changes you have:

(A X 2) $2S + 2O_2 \rightarrow 2SO_2$ $\Delta H = 2(-297\ kJ) = -594\ kJ$

Letting the desired reaction guide you, you see that SO_3 has to be a product, so equation (B) should be reversed. **When you reverse an equation, the sign of ΔH changes**. The reverse of equation (B) is the following:

(B reversed) $2SO_2 + O_2 \rightarrow 2SO_3$ $\Delta H = -198\ kJ$

Now add up the equations (A X 2) and (B reversed) to obtain the desired equation and add the ΔH values to determine the ΔH for the overall or net equation. **Any terms that are common to both sides of the combined equation should be canceled, just like an algebra problem.**

$$
\begin{array}{lll}
\text{(Ax2)} & 2S + 2O_2 \rightarrow 2SO_2 & \Delta H = 2(-297kJ) = -594kJ \\
\text{(B reversed)} & 2SO_2 + O_2 \rightarrow 2SO_3 & \Delta H = -198kJ \\
\hline
& 2S + 2SO_2 + 3O_2 \rightarrow 2SO_2 + 2SO_3 & \Delta H = -792kJ \\
& -2SO_2 \qquad\qquad -2SO_2 & \\
\hline
& 2S + 3O_2 \rightarrow 2SO_3 & \Delta H = -792kJ
\end{array}
$$

Solve the following Hess's Law problems:

1. Calculate the ΔH for the following reaction:

$Sn + 2Cl_2 \rightarrow SnCl_4$

You are given these two equations:

$Sn + Cl_2 \rightarrow SnCl_2 \qquad \Delta H = -325kJ$

$SnCl_2 + Cl_2 \rightarrow SnCl_4 \qquad \Delta H = -186kJ$

______ + ______ → _________ _______ = _______

______ + ______ → _________ _______ = _______

_______ + ______ + ______ → _______ + _______ _______ = _______

_______ + ______ → _______ _______ = _______

2. Calculate the ΔH for the following reaction:

$4Al + 3MnO_2 \rightarrow 2Al_2O_3 + 3Mn$

You are given these two equations:

$4Al + 3O_2 \rightarrow 2Al_2O_3 \qquad \Delta H = -3352kJ$

$Mn + O_2 \rightarrow MnO_2 \qquad \Delta H = -521kJ$

______ + ______ → _________ _______ = _______

_______ → _________ + _________ _______ = _______

_______ + ______ + ______ → _______ + _______ + _______ _______ = _______

_______ + ______ → _______ + _______ _______ = _______

3. Calculate the ΔH for the following reaction:

$2H_2O_2 \rightarrow 2H_2O + O_2$

You are given these two equations:

$2H_2 + O_2 \rightarrow 2H_2O \quad \Delta H = -572 \text{ kJ}$

$H_2 + O_2 \rightarrow H_2O_2 \quad \Delta H = -188 \text{ kJ}$

______ + ______ → _________ _______ = _______

_______ → _________ + _________ _______ = _______

_______ + ______ + ______ → _______ + _______ + _______ _______ = _______

____

______ → _______ + _______ _______ = _______

4. Calculate the ΔH for the following reaction:

$2CO + 2NO \rightarrow 2CO_2 + N_2$

You are given these two equations:

$2CO + O_2 \rightarrow 2CO_2 \quad \Delta H = -566.0 \text{ kJ}$

$N_2 + O_2 \rightarrow 2NO \quad \Delta H = 180.6 \text{ kJ}$

______ + ______ → _________ _______ = _______

_______ → _________ + _________ _______ = _______

_______ + ______ + ______ → _______ + _______ + _______ _______ = _______

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_______ + ______ → _______ + _______ _______ = _______

5. Calculate the ΔH for the following reaction:

$CS_2 + 2H_2O \rightarrow CO_2 + 2H_2S$

You are given these two equations:

$H_2S + 1.5O_2 \rightarrow H_2O + SO_2$ ΔH = -563 kJ

$CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2$ ΔH = -1075 kJ

______ + ______ → _________ + _______ _______ = _______

______ + ______ → _________ + _______ _______ = _______

_______ + ______ + ______ + ______ → _______ + _______ + ______ + ______ _______ = _______

______ + ______ → _________ + _______ _______ = _______

6. Calculate the ΔH for the following reaction:

$N_2 + 2O_2 \rightarrow 2NO_2$

You are given these two equations:

$N_2 + O_2 \rightarrow 2NO$ ΔH = 180.0 kJ

$2NO_2 \rightarrow 2NO + O_2$ ΔH = 112.0 kJ

______ + ______ → _________ _______ = _______

______ + ______ → _________ _______ = _______

_______ + ______ + ______ → _______ + _______ _______ = _______

_______ + ______ → _______ _______ = _______

Worksheet 5: Specific Heat

The original unit for heat was the calorie. The calorie was defined as the quantity of heat needed to raise the temperature of one gram of water by one Celsius degree. However, since 1960 with the adoption of the International System of Units (S.I.) **the unit for all forms of energy is the joule**. **Heat energy is therefore measured in joules or in calories. 1 calorie = 4.184 joules.** The calorie that is commonly used to measure food energy is written correctly as Calorie with a capital C and is actually 1000 calories or a kilocalorie. Most modern work uses joules as the unit for energy although calories are sometimes seen. Of course, **kilojoules (kJ)** (one thousand joules) **and kilocalories (kcal)** (one thousand calories) **are used for larger amounts of energy.**

Substances differ from one another in the amount of heat needed to raise their temperature. The specific heat of a substance is the heat required per gram per degree. The units for specific heat are Joules/gram ^{0}C. The specific heat for an element or a compound is constant (for the same state) and is useful information when studying the bonding, structure, and uses of the material.

Specific Heat of Water
1.00 cal/g ^{0}C or 4.184 J/g ^{0}C.

Questions:

1. The standard unit for energy is the __________________________.
2. _________________________ is the amount of heat to warm one gram of a substance one Celsius degree.
3. A container of lowfat yogurt contains about 150 Calories. What are the Calories really measuring?
4. How many joules are in 1 kilojoule?
5. When you get into a parked car in the summer time, the metal clip on the seatbelt is often too hot to touch, while the seatbelt fabric is not. What has a higher specific heat, the metal clip or the seatbelt fabric? Explain your answer using the definition of specific heat.
6. We could generalize the question above by saying that substances with a high specific heat get hot (faster/slower) than objects with a low specific heat.

For simple calorimetry problems we use the formula:
Heat = mass of sample X specific heat of material X change in temperature
Using symbols we can write this:
$H = m \times C_p \times \Delta T$

Usually books do not use signs for calculating whether the heat was absorbed or released but instead simply label the heat energy as joules absorbed or joules released depending upon whether the temperature went up or down.

__Example #1__: Calculate the number of joules needed to warm 100 grams of water from 25.0^{0}C to 80.0^{0}C.

$$\text{\# joules} = 100\,\cancel{g} \times 4.184\,\frac{J}{\cancel{g\,^0C}} \times (80.0\,^0\cancel{C} - 25.0\,^0\cancel{C}) = (100)(4.184)(55.0)\,J = 23{,}012\,J \text{ or } 2.30 \times 10^4\,J \text{ needed}$$

__Example #2__: Calculate the number of joules of heat released when 72.5 grams of water at 95.0^{0}C cools to 28.0^{0}C.

$$\text{\# joules} = 72.5\,\cancel{g} \times 4.184\,\frac{J}{\cancel{g\,^0C}} \times (95.0 - 28.0)\,^0\cancel{C} = (72.5)(4.184)(67.0)\,J = 20{,}323.78\,J \text{ or } 2.03 \times 10^4\,J \text{ released}$$

Solve the following problems on a separate sheet of paper. Be sure to label all numbers and show how the labels cancel out. All final answers should be circled. Round to two decimal places.

7. How many joules are needed to warm 25.5 grams of water from 14.0^0C to 22.5^0?

   Ans. 907 joules

8. Calculate the number of joules released when 75.0 grams of water are cooled from 100.0^0C to 27.5^0C.

   Ans. 2.28×10^4 joules

9. Calculate the heat, in joules, needed to warm 225 grams of water from 88.0^0C to its boiling point, 100.0^0C.

   Ans. 1.13×10^4 joules

10. **The specific heat of gold is 0.128 J/g^{0}C.** How much heat would be needed to warm 250.0 grams of gold from 25.0^0C to 100.0^0C?

    Ans. 2.40×10^3 joules

11. **The specific heat of zinc is 0.386 J/g^{0}C.** How many joules would be released when 454 grams of zinc at 96.0^0C where cooled to 28.0^0C?

    Ans. 1.19×10^4 joules

12. How many **calories** are needed to warm 15.0 grams of water at 12.0^0C to 86.0^0C?

    Ans. 1.11×10^3 cal

This basic equation can be solved for any of the four variables, heat (H), mass (m), specific heat (C_p) or change in temperature (ΔT)

Specific Heat (C_p) =	$\dfrac{\text{heat}}{\text{mass x change in temp}}$	or	$\mathbf{C_p = \dfrac{H}{m \times \Delta T}}$
Temperature Change (ΔT) =	$\dfrac{\text{heat}}{\text{mass x specific heat}}$	or	$\mathbf{\Delta T = \dfrac{H}{m \times C_p}}$
mass of material (g) =	$\dfrac{\text{heat}}{\text{change in temp x specific heat}}$	or	$\mathbf{m = \dfrac{H}{\Delta T \times C_p}}$

As you can see, all of three are simply using algebra to rearrange the formula at the top of the page. Notice how the units cancel so that only the desired unit remains.

Example #1: Calculate the **specific heat** of gold if it required 48.0 joules of heat to warm 25.0 grams of gold from 40.0^0C to 55.0^0C.

$$\frac{\text{\# joules}}{\text{gram } ^0\text{C}} = \frac{48.0 \text{ joules}}{25.0 \text{ g } (55.0 - 40.0)^0\text{C}} = \frac{48.0 \text{ joules}}{25.0 \text{ g} \times 15.0^0\text{C}} = \frac{0.128 \text{ J}}{\text{gram } ^0\text{C}}$$

Example #2: What **temperature change** will result from the addition of 3.22×10^3 joules of heat to 35.0 grams of water?

$$\text{Change in Temperature} = \frac{3.022 \times 10^3 \cancel{\text{joules}}}{35.0 \cancel{\text{g}} \times \frac{4.184 \cancel{\text{J}}}{\cancel{\text{g}}\, ^0\text{C}}} = 2.1988 \times 10^1 \, ^0\text{C} = 2.20 \times 10^1 \, ^0\text{C} \text{ or } 22.0^0\text{C}$$

Example #3: A sample of zinc, specific heat of 0.386 J/g^{0}C, released 1,964 joules of heat when it cooled from 92.5^0C to 65.0^0C. What was the **mass** of the zinc sample?

$$\text{\# g of Zn} = \frac{1964 \cancel{\text{Joules}}}{\frac{0.386 \cancel{\text{J}}}{\text{g}\cancel{^0\text{C}}} \times (92.5-65.0)\cancel{^0\text{C}}} = \frac{1964}{\frac{0.386}{\text{g}} \times 27.5} = 185.02 \text{ g} = 185 \text{ g Zn}$$

Example #4: What would be the **final temperature** if 8.94 x 10^3 joules of heat were added to 454 grams of copper, specific heat 0.386 J/g^{0}C, at 23.0^{0}C?

$$\Delta T = \frac{8.94 \times 10^3 \cancel{\text{joules}}}{454 \cancel{\text{g}} \times 0.386 \frac{\cancel{\text{joules}}}{\cancel{\text{g}}\ ^0\text{C}}} = 5.101 \times 10^1\ ^0\text{C} = 51.0^0\text{C} \rightarrow T_f = 23.0^0\text{C} + 51.0^0\text{C} = 74.0^0\text{C}$$

Solve the following problems neatly and orderly on a separate sheet of paper. Be sure to show all work, show how the units cancel, round to two decimals and circle the final answer.

13. How many joules are needed to warm 45.0 grams of water from 30.0^{0}C to 75.0^{0}C?

Ans. 8.47 x 10^3 joules

14. What would be the final temperature if 3.31 x 10^3 joules were added to 18.5 grams of water at 22.0^{0}C?

Ans. 64.8^{0}C

15. Calculate the specific heat of platinum if 1092 joules of heat were released when 125 grams of platinum cooled 65.2 Celsius degrees.

Ans. 0.134 J/g^{0}C

16. A sample of lead, specific heat 0.138 J/g^{0}C, released 1.20 x 10^3 joules when it cooled from 93.0^{0}C to 29.5^{0}C. What was the mass of this sample of lead?

Ans. 137 grams of Pb

17. When a 212 gram sample of iron at 93.5^{0}C was added to some cool water, the iron cooled to 36.5^0 and released 5.41 x 10^3 joules of heat. Calculate the specific heat of iron.

Ans. 0.448 J/g^{0}C

18. Calculate the final temperature is 2.10 x 10^3 joules are added to 325 grams of bismuth, specific heat 0.122 J/g^{0}C, at 27.0^{0}C.

Ans. 80.0^{0}C

19. How many joules are needed to warm 875 grams of iron, specific heat 0.448 J/g^{0}C, from 25.0^{0}C to 345^{0}C?

Ans. 1.25 x 10^5 joules

20. How many calories are released when 45.0 grams of water at 95.0^{0}C are cooled to 12.8^{0}C?

Ans. 3.70 x 10^3 cal

21. A sample of hot metal was poured into 112 grams of water at 25.5^{0}C. The heat from the metal warmed the water to a final temperature of 36.9^{0}C. How many joules were involved in warming this water?

Ans. 5.34 x 10^3 joules

Worksheet 6: Calorimetry

Measuring quantities of heat is called calorimetry. We can measure heat energy by measuring the temperature change it produces in a reference material, which is almost always water.

One of the major uses of calorimetry is to measure specific heats of metals. Specific heats are useful for engineers and manufacturers. This use of calorimetry is based upon the **law of conservation of energy – energy is neither created nor destroyed but can be transformed from one form to another.** In this application, **hot metal is added to cold water in an insulated container called a calorimeter. Heat flows from the hot metal to the cold water. Thermal equilibrium is when the two objects reach the same final temperature. The heat lost by the hot metal is gained by the cold water as they come to the same final temperature.** If you measure the temperature changes you can calculate the heat gained by the water and thus the heat lost by the metal, and finally can calculate the specific heat of the metal.

Questions:

1. When we say two objects are in thermal equilibrium we mean that their temperatures are ___________.
2. If I place a hot frying pan into a sink of cold water, what will happen to the temperature of the frying pan?

   What will happen to the temperature of the water?

Example #1: A 175 gram sample of a metal at 93.5^0C was added to 105 grams of water at 23.5^0C in a perfectly insulated container. The final temperature of the water and metal was 33.8^0C. Calculate the specific heat of the metal in J/g^0C.

Heat lost by the metal = heat gained by the water

Mass of Metal x specific heat of metal x temp change of metal = Mass of water x specific heat of water x temp change of water

$175\text{ g} \times C_p \times (93.5-33.8)^0C = 105\ \cancel{g} \times 4.184\ \frac{J}{\cancel{g^0C}} \times (33.8-23.5)\cancel{^0C}$

$175\text{ g} \times C_p \times 59.7^0C = 105 \times 4.184\text{ J} \times 10.3$

$C_p = \frac{(105)(4.184\text{ J})(10.3)}{(175\text{ g})(59.7^0C)} = 0.4331\ J/g^0C = 0.43\ J/g^0C$

Solve the following problems on a separate sheet of paper. You must use the system shown above, show all work, label all numbers, and show cancellation of labels, round to two decimals, and circle your final answer.

1. A 185 gram sample of copper at 98.0^0C was added to 102 grams of water at 20.0^0C in a perfectly insulated calorimeter. The final temperature of the copper-water mixture was 31.2^0C. Calculate the specific heat of copper using this data.

   Ans. 0.39 J/g^0C
2. A chemistry student added 225 grams of aluminum at 85.0^0C to 115 grams of water at 23.0^0C in a perfect calorimeter. The final temperature of the aluminum-water mixture was 41.4^0C. Use the student's data to calculate the specific heat of aluminum in joules/gram^0C.

   Ans. 0.90 J/g^0C
3. A student was given a sample of a silvery gray metal and told that it was either bismuth, specific heat 0.122 J/g^0C, or cadmium, specific heat 0.232 J/g^0C. The student measured out a 250 gram sample of the metal, heated it to 96.0^0C and then added it to 98.5 grams of water at 21.0^0C in a perfect calorimeter. The final temperature in the calorimeter was 30.3^0C. Use the student's data to calculate the specific heat of the metal sample and then identify the metal.

   Ans. 0.23 J/g^0C; cadmium

Worksheet 7: Mixed Calculations

This sheet contains a mixture of the calculations on Worksheet 5 (Specific Heat) and 6 (Calorimetry)

1. How many joules are needed to warm 675 grams of water from 12.0^0C to 85.0^0C?

   Ans. 2.06×10^5 J

2. How many calories are released when 35.0 grams of water at 100.0^0C are cooled to 0.0^0C?

   Ans. 3.50×10^3 cal

4. What was the mass of a sample of water if the addition of 4.28×10^3 joules raised the temperature of the sample from 22.0^0C to 34.5^0C?

   Ans. 81.8 grams

5. How many joules are needed to raise the temperature of 500. g of chromium, specific heat 0.448 J/g^{0}C, from 26.0^0C to 95.0^0C?

   Ans. 1.55×10^4 joules

6. What is the specific heat of a sample of an alloy if 3.75×10^3 joules of heat are released when a 315 gram sample of alloy cools from 78.0^0C to 28.4^0C?

   Ans. 0.240 J/g^{0}C

7. The specific heat of iron is 0.448 J/g^{0}C. What will be the final temperature if 3.50×10^4 joules of heat are added to a 454 gram sample of iron at 24.0^0C?

   Ans. 196^0C

8. A 245 gram sample of a metal at 99.5^0C was added to a 114 gram sample of water in a perfect calorimeter. The original temperature in the calorimeter was 23.5^0C. The final temperature of the metal-water mixture was 35.6^0C. Calculate the specific heat of this metal in joules per gram per Celsius degree.

   Ans. 0.369 J/g^{0}C

9. Calculate the number of joules needed to warm 275 grams of silver, specific heat 0.236 J/g^{0}C, from 25.0^0C to 950^0C.

   Ans. 6.00×10^4 Joules

10. What will be the final temperature when 8.75×10^3 joules are added to 75.0 grams of water at 23.0^0C?

    Ans. 50.9^0C

11. When 3.88×10^4 joules were added to a sample of iron, specific heat 0.448 J/g^{0}C, the temperature rose from 24.5^0C to 178^0C. What was the mass of the iron sample?

    Ans. 5.64×10^2 grams

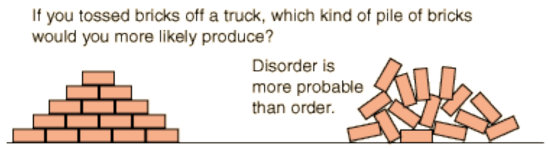
12. A 315 gram sample of tungsten at 92.5^0C was added to a 57.7 gram sample of water at 21.2^0C in a perfect calorimeter. The final temperature of the tungsten-water mixture was 31.8^0C. Use this data to calculate the specific heat of tungsten

    Ans. 0.134 J/g^{0}C

13. Calculate the final temperature that results from mixing 245 grams of cobalt, specific heat 0.446 J/g^{0}C, at 142^0C with 106 grams of water at 24.8^0C.

    Ans. 48.0^0C

Worksheet 8: Entropy and Gibbs Free Energy



In chemistry, the word spontaneous is used to describe reactions that will go forward without adding energy. Reactions in nature are driven by two forces (enthalpy and entropy), which together determine whether or not the reaction will be spontaneous. Reactions that are exothermic (release energy) are generally favored by nature and will be spontaneous. However, some endothermic reactions, such as the melting of ice, are spontaneous. The other thing that determines whether or not a reaction will occur is entropy. **Entropy can be defined as a measure of the degree of randomness of the particles, such as molecules, in a system. As might be expected from the chaotic world in which we live, nature favors an increase in entropy. In other words, reactions that increase the disorder of the system tend to be spontaneous.**

The amount of entropy of a system is best understood by considering the three principle states of matter. In a solid, the molecules vibrate in place and are not free to switch places with each other. **Solids are considered to have very low entropy, because very little randomness exists in them. Liquids are more disorderly than solids are therefore have higher entropy. Gases, the most disorderly of the three states possesses, have the highest entropy.** These are general guideline as some liquids (mercury, for instance) have lower entropy that certain solids. In general, dissolving a solid increases the entropy of a system.

Entropy of substances can be determined in the lab and are measured in molar values with units kJ/(mol·K). In this class, we will be most concerned with the change in entropy of a system, denoted ΔS.

Reactions in nature tend toward decreasing enthalpy and increasing entropy. A reaction that is exothermic and increases entropy will always be spontaneous. Conversely, an endothermic reaction that decreases the randomness of the system will never be spontaneous. However, what about reactions that decrease enthalpy but decrease entropy, or increase entropy but increase enthalpy? Will these reactions be spontaneous? The answer is that it depends on the temperature. In order to predict this, we must look at a factor called Gibbs Free Energy. **Gibbs free energy relates the enthalpy and entropy changes of a reaction:**

$$\Delta G = \Delta H - T\Delta S$$

If the Gibbs Free Energy of a reaction is negative, then the reaction will be spontaneous, if it is positive, then the reaction will not be spontaneous.

ΔH	ΔS	ΔG
Negative (exothermic)	Positive (more random)	Always negative (always spontaneous)
Negative (exothermic)	Negative (less random)	Negative (spontaneous) at lower temperature
Positive (endothermic)	Positive (more random)	Negative (spontaneous) at higher temperature
Positive (endothermic)	Negative (less random)	Never negative (never spontaneous)

1. $CH_3OH + 1.5\ O_2 \rightarrow CO_2 + 2\ H_2O$
   a. Methanol (CH_3OH) is used as rocket fuel. Methanol is liquid, while oxygen, carbon dioxide and water are all gases. Use this information to predict the sign of ΔS.
   b. Predict the sign of ΔH knowing that this is a combustion reaction.
   c. Does the sign of ΔG depend on temperature for this reaction?

2. Calculate ΔG using the Gibbs Free Energy equation. State whether or not the reaction will be spontaneous.
   a) $CH_3OH + 1.5\ O_2 \rightarrow CO_2 + 2\ H_2O$
   ΔH = -638.4 kJ ΔS = 156.9 J / K

   b) $2\ NO_2 \rightarrow N_2O_4$
   ΔH = - 57.2 kJ ΔS = -175.9 J / K