Thermochemistry

Heat and Energy

INFORMATION

Thermochemistry is the study of energy changes and transfers that occur during chemical reactions.

During chemical reactions, energy (in the form of heat) can either be consumed by the reaction (this is called an **endothermic** reaction) or heat can be released by the reaction (this is called an **exothermic** reaction). Energy stored in a compound is really stored in its chemical bonds – this is referred to as **chemical potential energy**. Sometimes, energy is required to break bonds; sometimes, it is required to form them.

Heat, represented by q, is a form of energy. Heat transfers from one object to another naturally as a result of the temperature difference between them. Heat *always* flows from warmer objects to colder ones. Heat will continue to flow until the temperature difference between them is zero.

Exothermic and Endothermic Processes

A **system** is a portion of the universe in which you have an interest. This may seem like somewhat of a daunting statement, but in reality, it simply indicates what you have your attention focused on for the purpose of the problem at hand. For example, a balloon filled with air can be considered a system; a beaker with two aqueous reactants can be a system; a refrigerator is a system; a gigantic chemical plant is a system. Determining what exactly constitutes a system is arbitrary – it is defined based on the needs or interests of the situation.

Surroundings, in contrast, represent everything else in the universe that is not in system. Again, this is broad and seemingly daunting, but usually the surrounding are defined as the area in the immediate vicinity (this may vary by scale – a reaction in a beaker may have "smaller" surroundings than a petrochemical factory, but you get the idea).

Thermochemistry focuses on the study of heat transfer (or **heat flow**) between the system and the surroundings. Exothermic processes result in heat leaving the system and entering the surroundings; endothermic processes result in heat leaving the surroundings and entering the system. Mathematically, heat entering a system has a positive sign ($+q$), and heat entering the surroundings has a negative sign ($-q$).

Recall that the **law of conservation of energy** states that energy can be neither created nor destroyed; therefore, the sum of the total heat in the system and the surroundings must remain the same. This means that as a certain amount of heat leaves the system, the same amount must enter the surroundings (and vice versa).

Heat, Quantitatively

The two common units used for measuring heat are the **calorie** and the **joule**. A calorie is defined as the amount of energy required to increase the temperature of one gram of pure water by one degree Celsius. This applies only to a calorie when written with a small "*c*" in its name. When written with a capital "*C*", as in **Calorie**, it refers to the more commonly known dietary calorie. Hence:

1000 calories (cal) = 1 kilocalorie (kcal) = 1 Calorie (Cal)

The energy denoted by a calorie, or by a Calorie, is measured by burning a sample under carefully monitored and controlled conditions and measuring the amount of heat released. When a candy bar wrapper indicates that the snack contains 100 Calories, this means that your body can break it down and acquire the equivalent of 100 Calories of energy. If one were to burn the same candy bar and measure the released heat, it would register as 100 Calories (or 100 kcal) of heat. This is, in fact, how Calorie content is determined for food.

The SI unit of energy is the joule. A joule is a smaller unit of energy than a calorie; one joule of heat can only raise the temperature of one gram of water by 0.239 degrees Celsius. Hence:

1 joule (J) = 0.0239 cal **or** 4.184 J = 1 cal

Heat Capacity

Water, iron, and polyethylene plastic all require different amounts of heat to warm them. This property of matter is called **heat capacity**, and is dependent on both the mass of a sample and its chemical composition. A larger sample of water requires more heat to increase its temperature than a smaller sample, for example.

Substance	c_p J/g·C
Asphalt	0.92
Brick	0.84
Concrete	0.88
Glass, silica	0.84
Glass, crown	0.67
Glass, flint	0.503
Glass, pyrex	0.753
Granite	0.79
Gypsum	1.09
Marble, mica	0.88
Sand	0.835
Soil	0.8
Wood	0.42

Early scientists observed that different substances require different quantities of energy to change their temperature. Have you ever noticed that the hood of a car is hot to the touch in the summer sun, but a swimming pool is cool when you jump in? This is because a given mass of a certain substance may require more or less heat than the same mass of another substance to facilitate a change in its temperature. This property of matter (the amount of heat required to raise the temperature of 1 gram of it by 1 degree Celsius) is called **specific heat capacity**, or just **specific heat**. Specific heat must be measured; it cannot be calculated from other observations; it is reported as **J/g·C**, read "joules per gram degree Celsius." Typically, specific heats are looked up on a table of common substances, like the one provided for common substances above. For reference, the specific heat of water is 4.184 J/g·C.

Heat and Energy Student Page Name______________________________ Hr_____

Questions:

1. What term is used when heat is ABSORBED by a reaction, thus leaving the surroundings colder?

2. What term is used when heat is RELEASED by a reaction, thus leaving the surroundings warmer?

3. Where is energy stored within a chemical?____________________

4. What symbol is used to represent heat? ___________

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5. Heat will always flow from a _______________________ object to a _____________________ object.

6. What do scientists call the item of interest or study?_____________________

7. Everything else but the item to be studied is called the __________________________.

8. The law of conservation of energy states that energy can be _________________ __________________ ___________________ ____________________

9. What are the two common units used for measuring heat?_____________ and _________________

10. A calorie is defined as ____

11. Calorie (note the capital letter) is defined as _______________________________________

12. The SI unit of energy is the ____________________________.

13. The heat capacity of an object is dependent on what two factors? __________________________ and ____.

14. Specific heat is a *measured* value should be reported with what units? ____________________

Part 2: Calorimetry

Calorimetry involves the measurement of heat flow from one object (often the system) to another (often the surroundings). We will assume that a calorimeter is a closed system where all the energy released/absorbed by an object is absorbed/released by the water in the calorimeter. If the mass of the water is known, the temperature change of the water can be used to determine the amount of heat energy released/absorbed. The simplest calorimeters consist of an insulated container, lid, stirring device, and thermometer or temperature probe. A sample is shown in figure 1 below:

Heat lost$_{object}$ = Heat gained$_{water}$

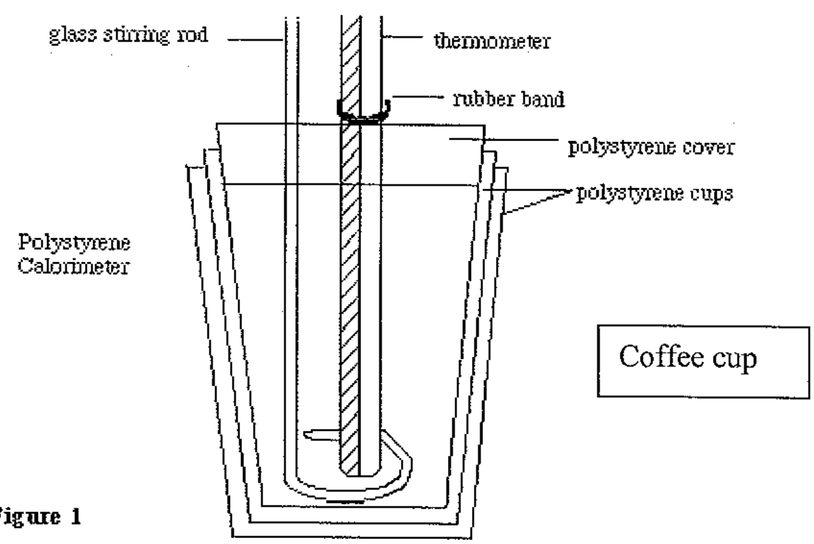

Figure 1

Equation for the calculation of heat is: **$q = mC\Delta T$**

q = heat released in Joules
m = mass of water in the calorimeter in grams
C = specific heat capacity (**not ^{0}C which is a temperature**)
ΔT = [final temperature – initial temperature]

Key Questions, Part 2:

1. What is the numerical value and units of the specific heat capacity of water?

2. What information does the specific heat capacity of water provide?

3. How can the heat released into some mass of water be calculated from the specific heat capacity of water and the change in temperature of the water (in other words, how is "q" calculated)? Answer in words not with an equation.

Critical Thinking Exercises, Part 2:

Answer Exercises 1-4 based upon the passage that follows.

A calorimeter was used to measure the heat released by when dissolving of a sample of sodium hydroxide. The calorimeter contained 100.0 g of water at an initial temperature of 10.0 ^{0}C. When the reaction was finished the temperature of the water increased to 75.0 ^{0}C. The specific heat capacity of water is 4.184 J/g ^{0}C

1. Write the mass of water (m) indicated in the passage.

2. Write the change in temperature (ΔT) indicated in the passage.

3. a) Write the correct mathematical set-up for the calculation of heat (q)?

   b) Substitute the appropriate values for m, C and ΔT in the equation and solve with the correct units included on work step labels and final answer.

4. If a substance with a *smaller* specific heat capacity than water were used in the experiment, identify whether ΔT would be larger or smaller. Explain.