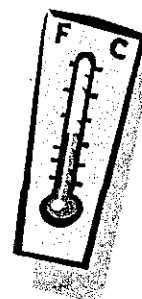
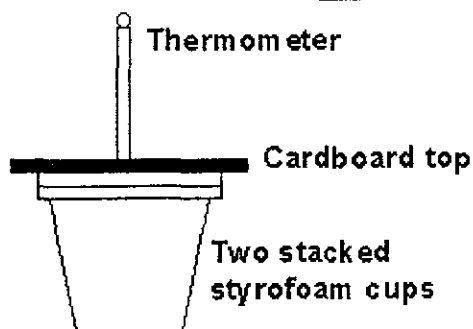


Activity: Calorimetry



Purpose:

To measure the flow of heat (ΔQ) from a hot metal cylinder into water, by using the formula $\Delta Q = mC\Delta T$.

Procedures:

1. Place a metal cylinder into a boiling water bath. Heat for at least 5 minutes.
2. Meanwhile, place approximately 100 mL of water in a Styrofoam cup. Record the EXACT volume measured. Record initial temperature.
3. When the metal is well heated, transfer it to the calorimeter. Cover IMMEDIATELY and place the thermometer through the lid.
4. Stir continuously. When the thermometer reading peaks, record the final temperature.
5. Clean up.

Data:

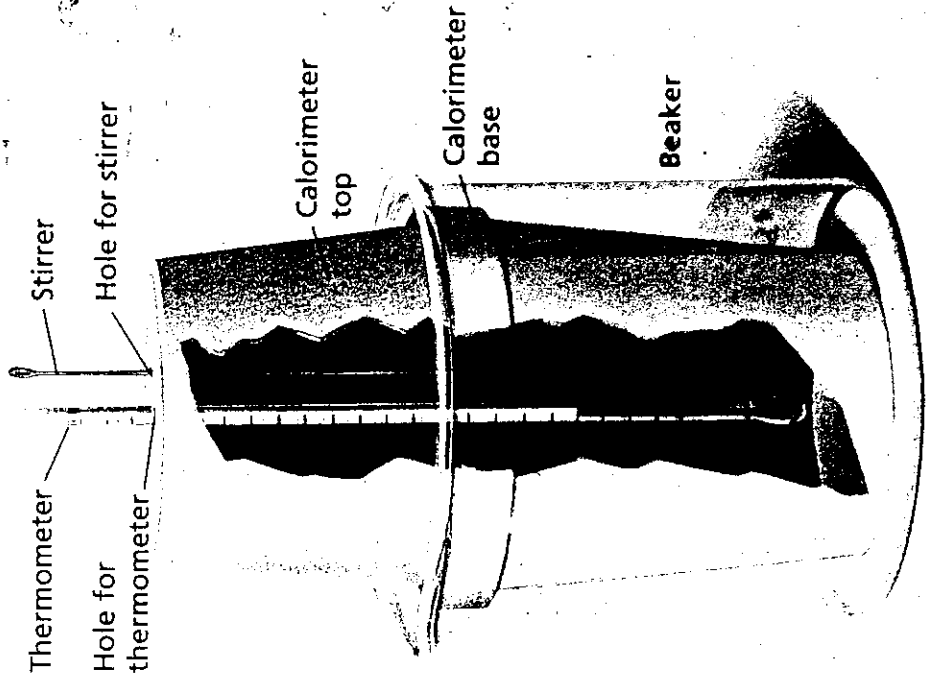
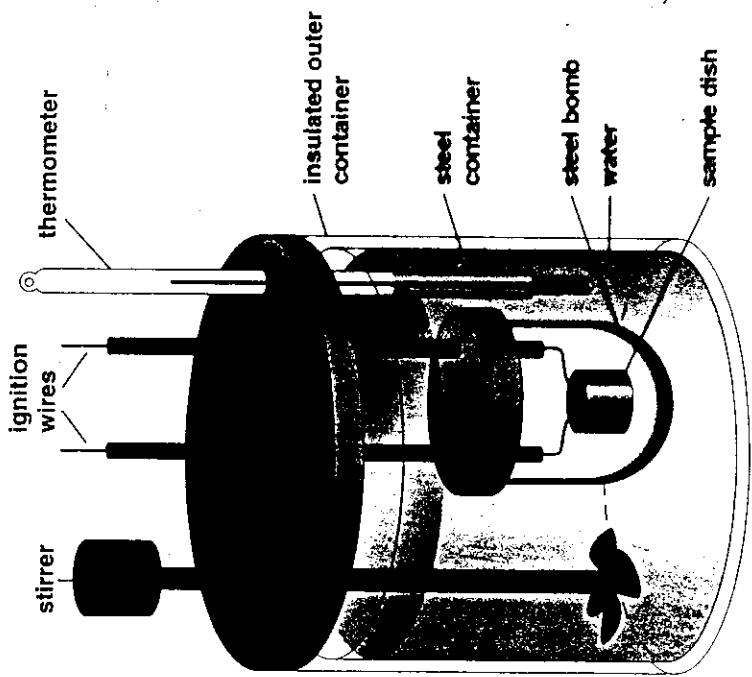
Volume of Water: _____

Initial Temperature: _____

Final Temperature: _____

Analysis Questions:

1. Knowing the volume of the water, what is the mass?
2. Calculate the ΔT for the water. (Show your work)
3. Calculate ΔQ for the water. (Show your work).
4. Describe the direction of heat in regards to the system and surrounding. Consider the metal to be the system and the water the surrounding.
5. Was the process exothermic or endothermic?
6. When the water reached its highest temperature, what temperature would you predict the metal was at? How do you know?



Non-SI Supplementary Problems

A non-SI unit often used to express the amount of heat exchanged in a reaction is the calorie (cal). The calorie is defined as the amount of heat needed to raise 1 g of water 1 degree Celsius. Therefore the specific heat of water is equal to 1.00 cal/g-°C. The formula used to calculate calories is:

$$\text{calories} = mc\Delta T$$

where m = mass of the substance heated

c = the specific heat of the substance

ΔT = the temperature change of the substance

Example: How much heat energy is required to raise the temperature of 5.00×10^2 g of water from 23.0°C to 28.4°C?

calories = mass of H₂O × specific heat H₂O × change in temperature of H₂O

$$\begin{aligned} \text{Change in temperature } (\Delta T) &= T_{\text{final}} - T_{\text{initial}} \\ &= 28.4^\circ\text{C} - 23.0^\circ\text{C} \\ &= 5.4^\circ\text{C} \end{aligned}$$

$$\begin{aligned} \text{heat} &= 5.00 \times 10^2 \text{ g} \times 1.00 \text{ cal/g}\cdot^\circ\text{C} \times 5.4^\circ\text{C} \\ &= 27 \times 10^2 \text{ cal} = 2.7 \times 10^3 \text{ cal} \end{aligned}$$

A larger unit of heat is the kilocalorie (kcal). One kilocalorie equals 1000 calories. The kilocalorie is also called the Calorie (Cal) or large calorie. This is the unit of measurement commonly used with foods.

Exercises

Solve the following problems on a separate sheet of paper, showing all work. Express your answers in the correct units with the appropriate number of significant figures.

- How many calories of heat are required to raise the temperature of 1.00 kg of water from 10.2°C to 26.8°C?
- How much heat is released when 275 grams of water cools from 85.2°C to 38.4°C? Express your answer in kcal.
- What temperature change will 100.0 mL of water undergo when it absorbs 325 calories of heat?
- What will the change in temperature be if 422 calories of heat is absorbed by 80.0 mL of water?
- What will the final temperature be if 45.0 mL of water at 15.4°C absorbs 2.50×10^2 calories of heat?
- What will the final temperature be if 688 calories of heat is released by 25 mL of water with an initial temperature of 80.0°C?
- A quantity of water is heated from 25.0°C to 36.4°C by absorbing 325 calories. What is the mass of the water?
- What is the mass of a sample of water that is heated from 10.0°C to 24.6°C while absorbing 1.00 kcal?
- What is the specific heat of lead if a 30.0 g of lead undergoes a 250°C change while absorbing 229.5 calories?
- A 1.00×10^3 -g block of aluminum releases 6.450×10^3 calories as it cools from 55.0°C to 25.0°C. What is the specific heat of aluminum?

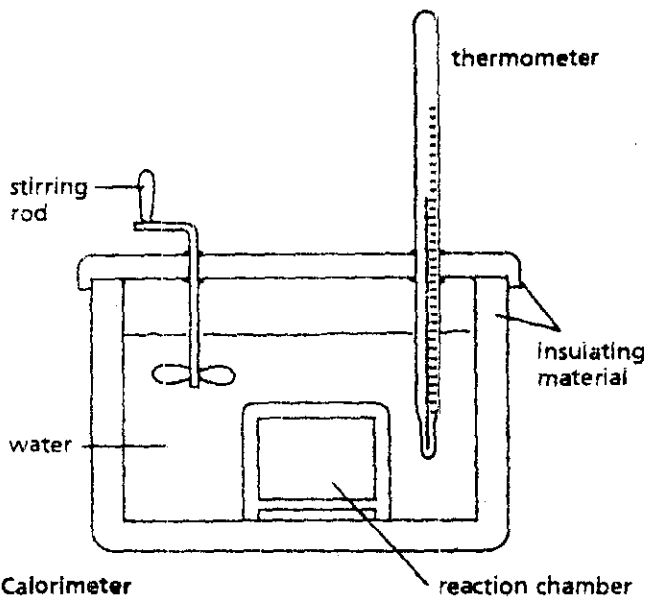
CHAPTER 5 REVIEW ACTIVITY

Text Reference: Section 5-6

Calorimetry

Answer the following questions.

- 1. a. The non-SI unit in which the quantity of heat energy has commonly been expressed is called the _____?
- 1. a. _____
- b. The SI-derived unit that is now the preferred unit for expressing this quantity is called the _____?
- b. _____
- 2. a. The quantity of heat energy required to raise the temperature of a unit mass of a substance by 1°C is called the _____? of the substance.
- 2. a. _____
- b. The value of this quantity for water, in SI-derived units, is _____?
- b. _____



- 3. a. In the calorimeter shown in the figure, a reaction that releases $8.4 \times 10^3 \text{ J}$ of heat energy takes place. Is this reaction endothermic or exothermic?
 - 3. a. _____
 - b. Assuming that the calorimeter contains $5.00 \times 10^2 \text{ g}$ of water and that the initial temperature is 30.0°C , what will the final temperature be? Choose your answer from among the following, and write the letter of the correct answer in the blank.
 - b. _____
- 1. 4.0°C
 - 2. 10.0°C
 - 3. 26.0°C
 - 4. 34.0°C
 - 5. 40.0°C

Activity: Using Calorimetry to Calculate Enthalpy

In your comp book, write pre-activity questions, data and observations and then answer analysis questions.

Pre-Activity Questions:

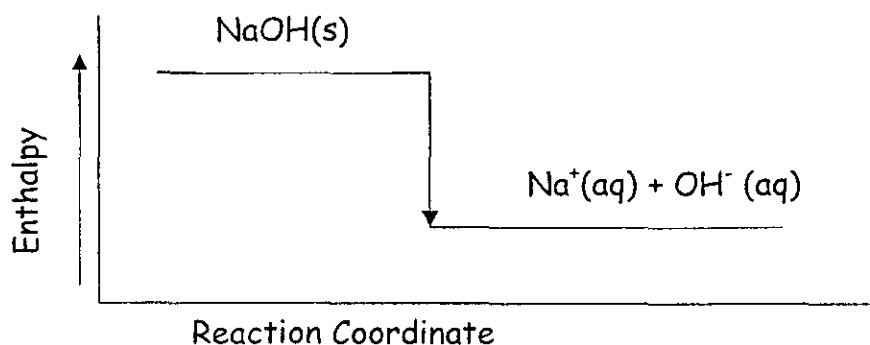
1. Write out the balanced equation for the dissociation of calcium chloride. Then write out the dissociation of ammonium nitrate.
2. Looking at your procedures, determine what is the mass of the water you will be using in the test tubes?
3. If the change in Q (surroundings) is positive, what will ΔH be? Will this reaction be exo or endothermic?
4. If the change in Q (surroundings) is negative, what will ΔH be? Will this reaction be exo or endothermic?

Procedure:

1. Place a scoop of calcium chloride or ammonium nitrate into a small plastic boat.
2. Add water to a test tube, measure the volume (about 10mL) and record.
3. Take the temperature of the water.
4. Add the salt and stir with the thermometer.
5. Take the final temperature
6. Repeat the steps with a different salt.

Analysis Questions:

1. Calculate the ΔQ for the water with calcium chloride (in Joules).
2. Calculate the ΔQ for the water with ammonium nitrate (in Joules).
3. If the salts are the system, what is the ΔH for each?
4. Label whether each reaction is exothermic or endothermic.
5. Draw a reaction diagram for each. (Look at the example below).
6. Which salt is MORE stable as a solid compound? Why?



- (Multiple choice) To determine the ΔH of a reaction, a calorimeter measures the energy changes of _____
 - The reaction
 - The system
 - The surroundings
 - The universe

- A calorimeter is filled with 75 mL of water at an initial temperature of 19°C . A 1.0-gram sample of NaOH is added to the calorimeter and the temperature increased to 26°C .
 - Calculate Q in joules.
 - Draw the calorimeter and use arrows to illustrate the heat transfer.
 - Is the reaction exothermic or endothermic?
 - Calculate ΔH for the reaction in joules.
 - Write the thermochemical equation (a balanced reaction including the energy term).

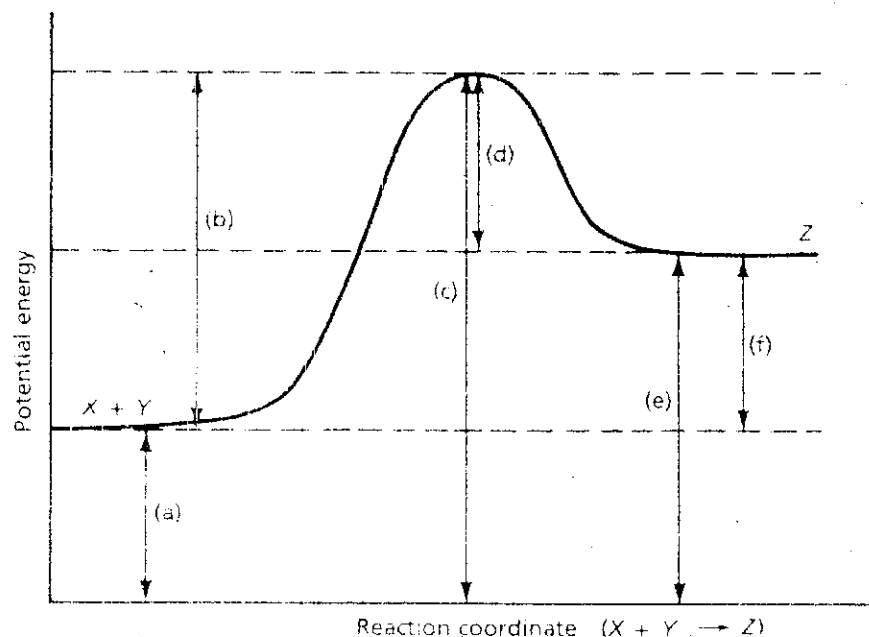
- A 4.25-gram sample of solid NH_4NO_3 is dissolved in 60 mL of water in a calorimeter. The temperature drops from 21°C to 16°C .
 - Calculate Q in joules.
 - Draw the calorimeter and use arrows to illustrate the heat transfer.
 - Is the reaction exothermic or endothermic?
 - Calculate ΔH for the reaction in joules.
 - Write the thermochemical equation.

CHAPTER 17 REVIEW ACTIVITY

Text Reference: Section 17-13

Potential Energy Diagrams

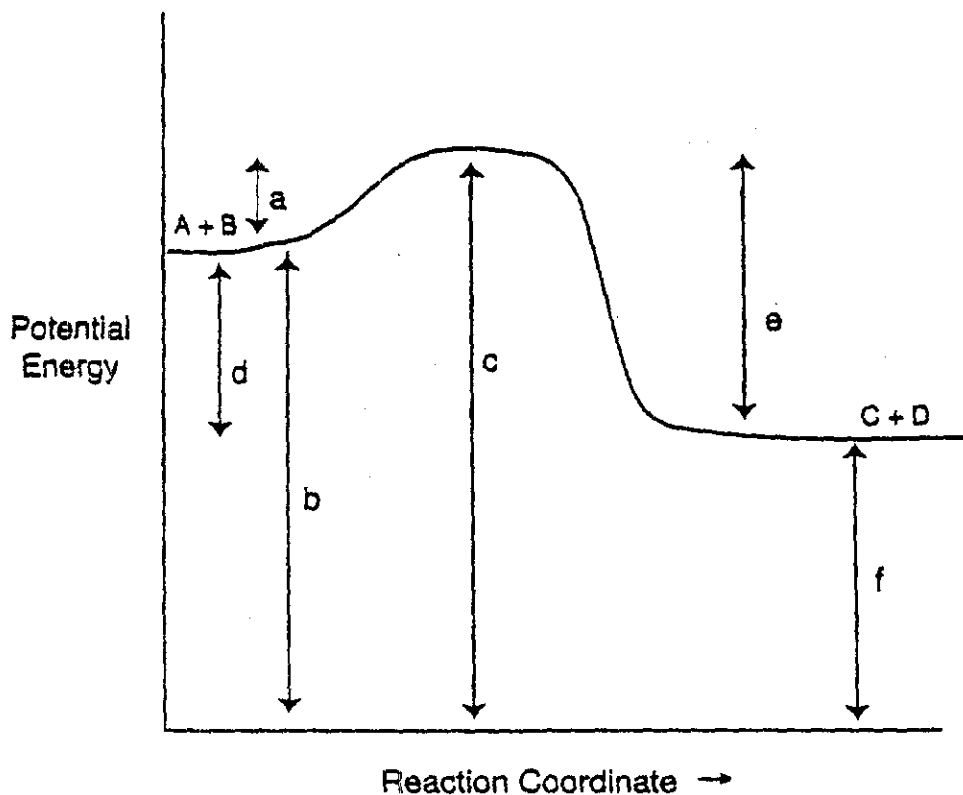
Answer the questions by referring to the diagram of the potential energy of a reaction.



1. Which of the letters (a)-(f) in the diagram represents the potential energy of the products? 1. _____
2. Which letter indicates the potential energy of the activated complex? 2. _____
3. Which letter indicates the potential energy of the reactants? 3. _____
4. Which letter indicates the activation energy? 4. _____
5. Which letter indicates the heat of reaction? 5. _____
6. Is the reaction exothermic or endothermic? 6. _____
7. Which letter indicates the activation energy of the reverse reaction? 7. _____
8. Which letter indicates the heat of reaction of the reverse reaction? 8. _____
9. Is the reverse reaction exothermic or endothermic? 9. _____
10. a. If a catalyst were added, which lettered quantities, if any, would change? 10. a. _____
 b. Would the activation energy increase, decrease, or remain unchanged? b. _____
 c. Would the heat of reaction increase, decrease, or remain unchanged? c. _____

POTENTIAL ENERGY DIAGRAM

Name _____

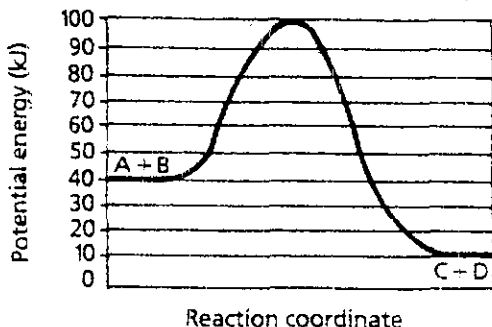
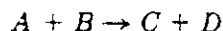


Answer the questions using the graph above.

1. Is the above reaction endothermic or exothermic? _____
2. What letter represents the potential energy of the reactants? _____
3. What letter represents the potential energy of the products? _____
4. What letter represents the heat of reaction (ΔH)? _____
5. What letter represents the activation energy of the forward reaction? _____
6. What letter represents the activation energy of the reverse reaction? _____
7. What letter represents the potential energy of the activated complex? _____
8. Is the reverse reaction endothermic or exothermic? _____
9. If a catalyst were added, what letter(s) would change? _____

Practice Problems

1. The graph below is a potential-energy diagram for the hypothetical reaction



a. Is the forward reaction endothermic or exothermic? Calculate the value of ΔH for this reaction.

b. Is the reverse reaction endothermic or exothermic? Calculate the value of ΔH for this reaction.

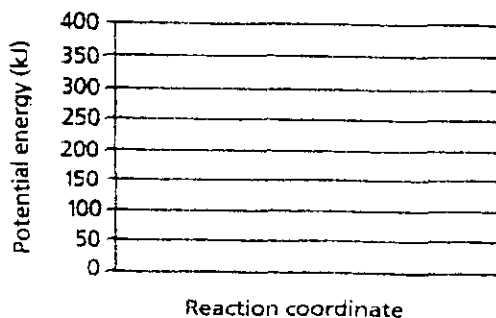
c. What is the value of the potential energy of the activated complex?

d. Calculate the activation energy for the forward reaction.

2. a. On the graph below, draw a potential-energy diagram for the following reaction



given the following assumptions: the potential energy of $Q + R$ is 150 kJ; the potential energy of $S + T$ is 250 kJ; the potential energy of the activated complex is 375 kJ.



b. Is the forward reaction endothermic or exothermic? Calculate the value of ΔH for this reaction.

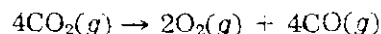
c. Calculate the activation energy for the forward reaction.

Refer to page 902 in the textbook when answering the questions below.

3. a. Calculate the amount of energy, in kilojoules, that is involved when 112.0 g of sodium chloride (NaCl) is produced from its elements. (Na = 23.0 g/mol; Cl = 35.5 g/mol)

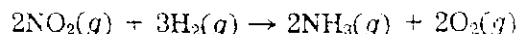
b. Is the energy absorbed or released? How can you tell?

4. a. Calculate the heat of reaction (ΔH), in kilojoules, for the reaction



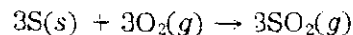
b. Is the reaction endothermic or exothermic? How can you tell?

5. a. Calculate the heat of reaction (ΔH), in kilojoules, for the reaction



b. Is the reaction endothermic or exothermic? How can you tell?

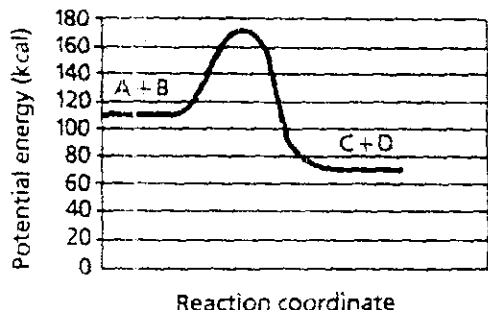
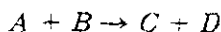
6. a. Calculate the heat of reaction (ΔH) for the reaction



b. Is the reaction endothermic or exothermic? Does the heat of reaction favor the forward reaction? How can you tell?

Non-SI Supplementary Problems

1. Consult the graph of a potential energy diagram for the reaction:



a. According to the graph what is the total amount of potential energy (in kilocalories) of the reactants?

b. How much potential energy must be added to the reactants in order to convert them into an activated complex? Is this part of the reaction endothermic or exothermic?

c. How much energy is released when the activated complex is converted into products? Is this part of the reaction endothermic or exothermic?

d. What is the net potential energy change between the reactants and the products? Is the overall reaction endothermic or exothermic?

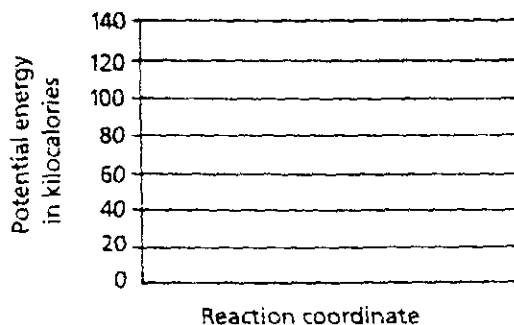
e. What is the activation energy of the reverse reaction?

f. Draw a dashed line to show the effect of adding a catalyst to the system

2. Given the equation:



Fill in the grid below using the following information:



potential energy of reactants: 72 kcal
 potential energy of products: 112 kcal
 activation energy for the forward reaction: 58 kcal

a. Is the forward reaction endothermic or exothermic?

b. What is the sign and value of the heat of reaction for the forward reaction?

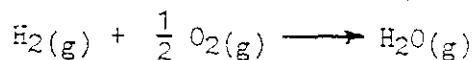
Consult the chart below for questions concerning heats of formation and free energy:

Compound	Reaction	Standard Heat of Formation (ΔH°) (kcal)	Standard Free Energy of Formation (ΔG°) (kcal)
calcium chloride	$\text{Ca}(s) + \text{Cl}_2(g) \rightarrow \text{CaCl}_2(g)$	-190.0	-179.3
carbon dioxide	$\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g)$	-94.1	-94.3
carbon monoxide	$\text{C}(s) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{CO}(g)$	-26.4	-32.8
ethyne	$2\text{C}(s) + \text{H}_2(g) \rightarrow \text{C}_2\text{H}_2(g)$	53.8	50.0
hydrogen sulfide	$\text{H}_2(g) + \text{S}(s) \rightarrow \text{H}_2\text{S}(g)$	-4.80	-7.89
potassium chloride	$\text{K}(s) + \frac{1}{2}\text{Cl}_2(g) \rightarrow \text{KCl}(s)$	-105.0	-97.6
methane	$\text{C}(s) + 2\text{H}_2(g) \rightarrow \text{CH}_4(g)$	-14.9	-12.1
silver bromide	$\text{Ag}(s) + \frac{1}{2}\text{Br}_2(l) \rightarrow \text{AgBr}(s)$	-23.8	-22.9



Activity 12-6 Practice Exercise

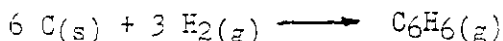
1. The equation for the formation of water can be written as



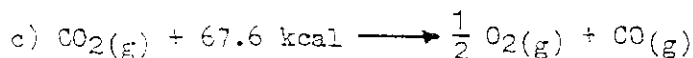
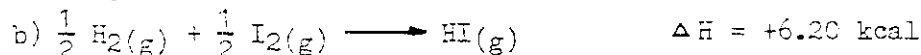
- a) The reaction liberates 57.83 kcal of heat to the surroundings for every mole of H_2O formed. Based on this information, write a thermochemical equation for this reaction (add an energy term to the appropriate side of the equation).
- b) For this reaction, $\Delta H = \underline{\quad ? \quad}$.

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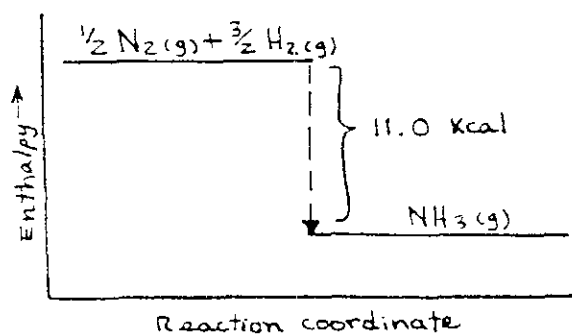
2. The equation for the formation of benzene is



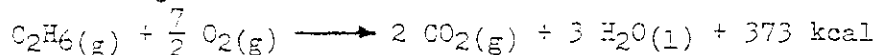
- a) This reaction absorbs 19.82 kcal of energy from the surroundings for every mole of C_6H_6 formed. In light of this information, write a thermochemical equation for the reaction.
- b) For this reaction, $\Delta H = \underline{\quad ? \quad}$.
3. State whether each of the following is exothermic or endothermic.



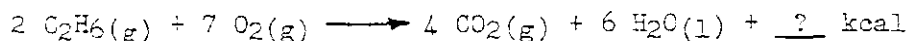
4. Write a thermochemical equation for the reaction shown in the reaction diagram at right.



5. Ethane burns according to

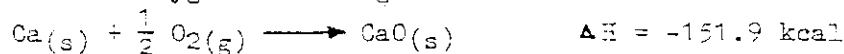


What is the energy term for the reaction when it is written as



(Be sure to note how the coefficients compare in the two equations.)

6. Calcium combines with oxygen according to



How many kcal of energy will be released when 0.5 mole of $\text{Ca}(\text{s})$ is completely reacted with 0.25 mole of $\text{O}_2(\text{g})$? (Hint: the amount of energy released is proportional to the amounts of reactants consumed.)

Answers are given at the end of the chapter.

Review Thermochemistry

1. In the space provided below, plot the potential energy diagram based on the following information:
 - a) potential energy of the reactants = 200 kJ
 - b) potential energy of the products = 275 kJ
 - c) potential energy of the activated complex = 350 kJ

For questions # 2-4, refer to the above potential energy diagram you plotted.

2. What is the activation energy (E_A) for the above reaction?
3. Using the diagram above, calculate the ΔH for the reaction.
4. Is the reaction endothermic or exothermic?
5. Using a calorimeter, the ΔQ for a reaction is -250 calories. If the calorimeter has 400 mL of water that has a starting (or initial) temperature of 20 °C, what will the final temperature be ?
6. Sketch a potential energy diagram for an exothermic and endothermic reaction.

12-4 Apply

The Human Body as an Engine

Use the chart below to answer the questions that follow.

Food (Portions)	Calories	Time (Minutes) Spent				
		Lying Down, Resting	Walking	Bicycle Riding	Swimming	Running
Apple (large)	101	78	19	12	9	5
Bacon (2 strips)	96	74	18	12	9	5
Carrot, raw	42	32	8	5	4	2
Chicken, fried ($\frac{1}{2}$ breast)	232	178	45	28	21	12
Egg, fried	110	85	21	13	10	6
Halibut steak ($\frac{1}{4}$ pound)	205	158	39	25	18	11
Hamburger sandwich	350	269	67	43	31	18
Malted milk shake	502	386	97	61	45	26
Milk (1 cup)	150	115	28	19	13	7
Orange juice (1 glass)	120	92	23	15	11	6
Pizza, cheese ($\frac{1}{8}$)	180	138	35	22	16	9
Potato chips (1 serving)	108	83	21	13	10	6
Soda (1 glass)	106	82	20	13	9	5
Tuna fish salad sandwich	278	214	53	34	25	14
Shrimp, fried (1 piece)	180	138	35	22	16	9

Note: 1 Calorie = 1000 calories, or 1 kilocalorie.

Questions

- How many Calories are there in half of a cheese pizza?

- How many minutes of bicycle riding would be required to utilize the Calories in half of a cheese pizza?

- Calculate the approximate number of Calories that are burned per minute when resting, walking, bicycle riding, swimming, and running.

- How long would a person have to walk in order to burn off the Calories in a meal consisting of a glass of soda, a hamburger sandwich, and a serving of potato chips? How long would it take to burn off the same meal if the person rode a bicycle? What if the person went running?

12-4 Apply (continued)

5. Why is it important not to consume an excess of Calories on a daily basis?

6. The calories in food come from the burning of carbohydrates, fats, and proteins. Carbohydrates provide 4 Cal/g and fats provide 9 Cal/g. A glass of milk, which has 150 Cal, has 8 g of fat and 11 g of carbohydrate. Calculate the number of Calories that come from fats and from carbohydrates. What percentage of the total Calories in milk come from fats? From carbohydrates?

7. A meal consisting of a hamburger sandwich, a serving of potato chips, and a glass of soda, contains 67 g of carbohydrates and 21 g of fat. Calculate the percentage of the total Calories in the meal that come from carbohydrates and the percentage that come from fats. (Hint: Use the chart to calculate that total number of Calories in the meal.)

8. Basal metabolic rate (BMR) is the measurement of the rate at which the body expends energy under normal resting conditions. For the average 70-kg person, the BMR is 72 Calories/hour. This represents the minimum amount of energy required to sustain life functions. Calculate the minimum required Calories for a 70-kg person in a day.

Fast Fat

by Kim D. Benson

Fast food restaurants have a special aroma. One whiff can stoke your appetite and trigger mental images of juicy hamburgers and hot French fries. That aroma is the smell of fat. If we give in to its appeal too often, we may end up overweight and have a greater chance of having a heart attack.

Humans are the only animals that cook meat, so we are the only species to be tempted by its aroma. It may be just a coincidence, but humans also carry more fat than most animals. Among the findings of a recent study by Caroline Pond and Christine Mattacks of the Open University in England are the following: carnivores are more naturally fat than herbivores; domestic animals are fatter than their wild counterparts; and humans have at least ten times more fat cells in proportion to body mass than most other mammals. This last finding may change the traditional view of how fat humans should be. Some of the earlier conclusions were based on extrapolations from the fat content of laboratory rats and mice, which are at the lower end of the mammal fat scale.

Fat chance

Fat may be part of our natural condition, but now that we are living far longer than our ancestors, fat has become an enemy. According to the National Academy of Sciences, a diet high in fat has been repeatedly connected with cancer of the breast, prostate, and bowel. Fat makes up about 40% of the Calories in the average American diet, but nutritionists suggest that fat should be no more than 30% of our total Calories.

(One nutrition Calorie is equivalent to 1,000 scientific calories, or 1 kcal. The nutrition Calorie is often capitalized to emphasize this distinction.) For someone eating 2,000 Calories per day, that's about 15 teaspoons of fat per day. In 1982, former Secretary of Health and Human Services Margaret Heckler estimated that 30% of cancer cases could be avoided by an increase in fiber consumption and a reduction in fat intake. Yet a small amount of fat is necessary to the diet because it is natural source of vitamins A, D, E, and K, and it contains polyunsaturated fatty acids essential for maintaining growth and good health.

Fats also supply us with long-lasting energy. One gram of fat will supply nine Calories of energy, whereas one gram of protein or carbohydrate supplies only four Calories. That means nine Calories of energy are required to burn one gram of fat, but only four Calories are needed to burn one gram of protein or carbohydrate.

Fats not only provide more energy than other foods, they also remain in the stomach longer and provide much of the satisfied sensation you feel after a meal. Perhaps that is why we enjoy eating fatty foods to excess. If we had enough drinking water, we could probably survive for five weeks on the fat stored in our body.

Fats and oils are an integral part of the way we cook. The brown in hash browns comes from the oil they are cooked in. Fat gives cookies their light and crispy texture and ice cream its creamy smoothness.

What's the difference between a fat and an oil? Both are triglycerides, whose molecules are made from three fatty acids. Triglycerides that are solid at room temperature are called oils. Since fats and oils have the same energy content, food labels lump them together and call them both FAT.

Pertinent percent

Check the label on your favorite breakfast cereal. You'll see that it gives the grams of fat, protein, and carbohydrates in one serving. The label clearly lists the grams of fat in the cereal, but it is harder to determine how many Calories of fat it contains. To do this, you'll need to find the total Calories and grams of fat in one serving. Convert the grams of fat into Calories:

$$\text{Grams fat} \times 9 \text{ Cal / g} = \text{Calories from fat}$$

Then find what percentage this is of the total Calories:

$$\text{Calories from fat} \times 100 / \text{total Calories per serving} = \% \text{ Calories from fat}$$

This value, the percentage of Calories that come from fat, is one of the most important ways to evaluate the fat content of your food, and it holds some surprises. The content of fat in foods such as butter and cooking oil (both 100% fat) is obvious, but hot dogs can deliver more than 80% of their Calories as fat, and potato chips can be over 50%.

The other important factor is whether or not the fat is saturated. Nutritionists commend that, of our daily fat allowance (30% of total Calories), no more than 10% should come from saturated fat. Saturated fat is found in dairy products, coconut and palm oil, and meat. The higher the grade of meat (prime), the more "marbling" or saturated fat it will contain. The differences between saturated and unsaturated fat are easy to see when the molecules are drawn on paper (see below). Identifying them when they are sitting on a dinner plate is a bit more difficult.

Starts at home

To adjust our fat intake when we cook at home, we can use less fat in general. Remember, a high-fat diet increases the risk of colon cancer and can make you overweight. When you do use fat, you should try to replace animal fat with poly- or monounsaturated vegetable oil. For example, don't fry eggs in bacon grease; use unsaturated oil instead. There are many cooking oils available, and they differ greatly in their content of saturated fats. Safflower, olive, canola, corn, soybean, and some others are desirable because of their low saturated fat content. If you love fried chicken, don't cook it in lard or shortening; use soybean, olive, or corn oil instead. It'll be just as tasty and healthier. You would never consider spreading beef fat (31% saturated) on a slice of bread, but you may think nothing of a generous application of butter, which is more saturated. If you fry food in butter, you probably get more saturated fat than you do from the hamburger at McDonald's.

We can control the amount and type of fat we eat at home, but when we eat out, it's more difficult to keep control. Take French fries, for example. French fries are deep fried-immersed in hot fat. In some cases the fat used is tallow (beef fat) because it adds an appealing flavor to an otherwise bland piece of potato. Unfortunately, tallow is 31% saturated fat. Many fast food restaurants deep fry their chicken and fish in vegetable oil, but many restaurants don't say which vegetable oil they use. If it's soybean oil, you'll benefit from getting a relatively small amount of saturated fat. However, palm kernel oil and coconut oil, which are also used, are more saturated than tallow. If a restaurant won't tell you which kind of vegetable oil it uses, consider eating somewhere else.

So how do we avoid consuming excess fat? We can read labels, cook more from scratch (rather than using package mixes) so that we are in control of the quantity and type of fat, and we can choose foods and snacks that are lower in fat, especially saturated fats. Put away the chocolate chip cookies and bring out the raisins!

Measuring Calories

Energy Content of Food

Purpose:

To use calorimetry to find the nutritional calories in snack food and compare that to the nutrition label on the snack food package.

Pre-Lab Questions:

- 1) A candy bar has a total mass of 2.5 ounces. In a calorimetry experiment, a 1.0g sample of this candy bar was burned in a calorimeter surrounded by 1000g of water. The temperature of the water in contact with the burning candy bar was measured and found to increase from an initial temperature of 21.2 °C to a final temperature of 24.3 °C.
 - a. Calculate the amount of heat in *calories* released when the 1.0g sample burned.
 - b. Convert the heat in calories to nutritional Calories (1000 calories = 1 Calorie) and divide by the mass of the burned sample in grams to obtain the *energy content* (also called fuel value) in units of Calorie per gram.
 - c. Multiply this value by the total number of grams in the candy bar to calculate the total *calorie content* of the candy bar in Calories. *Hint: convert the mass in ounces to grams. (1 oz. = 28.38 grams)*
- 2) Consult the nutritional labels on two of your favorite snack foods: Report their total calorie content (total Calories) and calculate their fuel value (Calories per gram).

Materials:

- Balance, Centigram (.01g precision)
- Soda Can (calorimeter and lid)
- Erlenmeyer flask with plastic spill-rim collar, 125 mL
- Food holder (cork) and pin
- Graduated Cylinder, 50mL
- Matches
- Snack foods, 2 pieces
- Stirring rod
- Thermometer or temperature sensor
- Water

Procedures:

- 1) Place a food sample on the food holder. Measure and record the combined mass of the food holder and sample. Place the food holder on a ring stand
- 2) Obtain a clean, empty soda can.
- 3) Obtain about 50mL of tap water. Record the EXACT volume of the water. Add the water to the empty soda can.

- 4) Bend the top tab on the can up and slide a stirring rod through the hole. Suspend the can on a ring stand using a metal ring. Adjust the height of the can so that it is about 2.5cm above the food holder.
- 5) Insert a thermometer into the can. Measure and record the initial temperature of the water.
- 6) Light the food sample and center it under the soda can. Allow the water to be heated until the food sample stops burning. Record the maximum (final) temperature of the water in the can.
- 7) Measure and record the final mass of the food holder and sample.

Data Table: *The Calorimetry Experiment*

Volume of water (in mL)	Initial Mass of Food Sample and Holder	Final Mass of Food Sample and Holder	Initial Temperature of Water	Final Temperature of Water
*Mass is of Food Sample and Holder (in grams)				
*Temperature is of water (in degrees C)				

Post-Lab Calculations and Analysis

Construct a *Results Table* to summarize your results from all of the following calculations.

- 1) Determine the mass of water heated in the calorimeter for each food sample.
- 2) Calculate the change in temperature for each sample.
- 3) Use the heat equation to calculate the ΔQ absorbed by the water in the calorimeter for each food sample. Report the results in *calories and nutritional Calories*.
- 4) Subtract the final mass of the food sample and holder from the initial mass to determine the mass in grams of the food sample that burned in each experiment.
- 5) Use the results from Questions #3 and #4 to calculate the energy content (fuel value) of the food sample in units of *Calories per gram*. (Cal/g) Record this value.
- 6) Compare your results with the label on the package. Calculate % error. Place in your results table.
- 7) Consider the major sources of error in this experiment. Do you think your results are off on the high side or the low side? Explain.