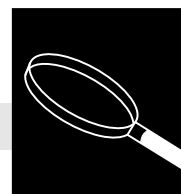




# Topic Overview



## CONTENT IN A NUTSHELL

You feel cold before you dry off after a swim or shower because each mole of water on your skin requires approximately 40 kJ (10 kcal) of heat to evaporate, and this heat is taken from your body. The chemical cold packs used to treat sports injuries cool their surroundings because the chemical reaction absorbs heat from the water when the chemicals in the packs dissolve (go into solution). In contrast, concentrated sulfuric acid releases large quantities of heat when added to water. Chemical hot packs provide heat in much the same way—except that they do not contain sulfuric acid!

You are kept warm in winter by the heat released when fuels such as coal, oil, or natural gas are burned. Heat released from nuclear fission reactions in reactor cores is used to generate electricity by nuclear power plants, while the sun's energy comes from the nuclear fusion reactions deep in its interior. Our Earth receives only 1 part in 2.4 billion parts of the sun's energy ( $1.7 \times 10^{17}$  watts out of  $4 \times 10^{26}$  watts). It would require 115 million nuclear (fission) power plants to generate this same quantity of energy. It takes only three days for the Earth to receive from the sun a quantity of energy equivalent to all fossil fuel reserves on Earth. The United States, which has over 40% of the world's coal reserves, burns about 700 million tons of coal each year to generate over half of our needed electrical energy. It is estimated that, by the year 2005, it will take as much energy to produce a barrel of oil as we would obtain by burning that quantity of oil as a fuel. Future energy sources include solar energy, biomass, wind power, hydroelectric, geothermal, tidal, fuel cells, and nuclear fission and fusion reactions.

Different substances have different quantities of energy stored in them as potential energy. When substances undergo physical or chemical change, energy is released or absorbed as heat and other forms of energy. Energy is conserved during such changes, which allows chemists to predict the quantity of energy released or absorbed. The quantity of energy released or absorbed can be measured indirectly using the technique known as calorimetry, which stems from the old definition of the unit of heat energy known as the calorie. Since one calorie (cal) is the quantity of heat needed to raise the temperature of one gram of water just one degree Celsius, calorimetry involves measuring the temperature change of a known mass of water accompanying the studied change. The IUPAC unit of heat energy, the joule (J), is the preferred unit. 1 cal equals 4.184 J, or 1 J equals 0.2390 cal.

If a water calorimeter is used and the water temperature drops during the change being measured, heat energy must have been absorbed from the water, and the change is said to be **endothermic**. If water temperature rises, heat was released to the water, and the change is **exothermic**.

Whenever chemical equations representing two or more changes can be algebraically combined to yield a new equation, the algebraic sum of the heats accompanying these changes is found to equal the quantity of heat associated with the net change. This regularity, known as **Hess's Law**, permits calculation of some heats of reaction without measuring them directly.

A table of standard heats of formation of compounds from their elements provides another tool students can use to make the same calculated prediction of heats of reaction.

Thermochemistry fits into the curriculum in several places. It is related to stoichiometry (heats of reaction), chemical bonding (bond energies), phase changes (heats of vaporization and fusion), solution chemistry (heats of solution and hydration), and nuclear chemistry (fission and fusion). It should be taught prior to the discussion of rates of reaction, since activation energies for endothermic and exothermic reactions are related to both chemical kinetics and equilibrium.

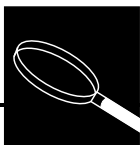
## PLACE IN THE CURRICULUM

## CENTRAL CONCEPTS

1. A **system** is that portion of the physical universe under study. Other portions are termed the **surroundings** or the environment.
2. **Heat** is a form of energy that flows across a conducting boundary from the system at a higher temperature to the system at a lower temperature (from a hot body to a cold body).
3. When chemical and/or physical changes occur, they involve definite amounts of matter accompanied by specific quantities of heat energy.
4. **Endothermic** processes are those in which heat is absorbed by the system from the surroundings; **exothermic** processes release heat from the system to the surroundings. Processes in which no release or absorption of heat occurs are termed **adiabatic** changes.
5. A **joule** (J) is the energy that a 1-kg mass gains when it is acted upon by a force of 1 N (newton) through a distance of 1 m. The older unit, the calorie (cal), is equivalent to 4.184 J.
6. **Heat capacity** of a material is defined as the quantity of heat required to raise the temperature of a substance by one Kelvin. ( $1\text{ }^{\circ}\text{C} = 1\text{ K}$ ) The specific heat capacity is the heat capacity per gram of substance. Common examples of specific heat capacity ( $\text{J K}^{-1}\text{g}^{-1}$ ) are: water = 4.18; copper = 0.38; nitrogen = 1.04; and ammonia = 2.06. The molar heat capacity is the heat capacity per mole of substance.
7. **Calorimetry** is an experimental procedure in which the quantity of heat released or absorbed during a change is indirectly measured, at constant volume or at constant pressure, for a known quantity of matter. From the definition of a calorie, it follows that this quantity of heat for a water calorimeter can be calculated using the relation

$$\begin{array}{cccc} \text{Quantity heat} & = & \text{Mass water} & \times & \text{Temp. change water} & \times & \text{Specific heat water} \\ \text{J (or cal)} & & \text{(g)} & & \text{(K or } ^{\circ}\text{C)} & & \text{(JK}^{-1}\text{g}^{-1}\text{ or cal/g } ^{\circ}\text{C)} \\ Q & & m & & \Delta T & & C \\ Q & = & m & \times & \Delta T & \times & C \end{array}$$

8. The heat involved in a physical or chemical change occurring at constant pressure is known as the **enthalpy change**, given the symbol  $\Delta H$ .
9. Tables listing the quantities of heat accompanying many chemical changes have been compiled, in units of both joules and calories. When such changes involve the formation of one mole of a chemical compound from its elements in their normal (standard) states, the values are known as standard molar heats of formation,  $\Delta H_f^{\circ}$ .



10. The heat energy term associated with a chemical or phase change can be included in the corresponding equation. Heat can be treated as a product or a reactant in an equation. If heat is absorbed during the change, the heat term is treated as a reactant, and placed to the left of the arrow in the equation. If heat is released, this term is treated as a product, and is placed to the right of the arrow, with the substances produced in the reaction. Alternatively, the heat energy term can be listed apart from the equation, at the end, as a  $\Delta H$  notation, where a + sign indicates an endothermic reaction and a - sign indicates an exothermic reaction.
11. The first law of thermodynamics sets down the empirical principle that the internal energy change ( $\Delta E$ ) in a system equals the sum of the heat ( $q$ ) and work ( $w$ ) in that system ( $\Delta E = q + w$ ). The internal energy is the sum of the translational, rotational, vibrational, electronic, nuclear, and zero point energies ( $E = E_{\text{trans}} + E_{\text{rot}} + E_{\text{vib}} + E_{\text{elec}} + E_{\text{nuclear}} + E_{\text{zero point}}$ ).
12. The total energy of a system is composed of kinetic energy and potential energy. **Kinetic energy**, energy of motion, is directly related to temperature. The higher the temperature, the greater the kinetic energy of the system. **Potential energy**, energy of position, is associated with attractions and repulsions among (and within) the molecules, atoms or ions present in the system. Different substances are assumed to contain different quantities of stored energy (potential energy). When one substance changes to another, with a new potential energy content, energy is either released or absorbed as heat energy. The total kinetic energy of a polyatomic system consists of three types: translational, rotational, and vibrational. Solids exhibit vibrational energy only, while liquids and gases can exhibit all three modes of motion.
13. Phase changes, such as melting and evaporation, involve energies of 2-10 kJ/mol. Chemical changes involve ten to a hundred times more energy than phase changes. Nuclear changes involve energies millions of times greater than chemical changes.
14. The assumption that total energy is conserved during a chemical or physical change permits the calculation of the change in stored energy, also known as the change in heat content ( $\Delta H$ ), when the change is measured in a calorimeter. If one mole of A absorbs 10,000 J when changing to one mole of B, for example,  $\Delta H = +10$  kJ per mole of B formed. One mole of B would have 10 kJ more potential energy than did the one mole of A. If 10,000 J had been released in the process, then  $\Delta H = -10$  kJ/mol of B. Mole for mole, in this latter case, B would have 10 kJ less potential energy than A.
15. Whenever equations representing two or more changes can be combined algebraically to give the equation of another, different change, the algebraic sum of the heats accompanying these changes is found to equal the quantity of heat associated with the net change. This is known as **Hess's Law**.
16. From the concept that heat (enthalpy) of reaction is  $\Delta H$  (where "delta" means a change in some fundamental property) one can calculate the heat of reaction for a given equation using standard molar heats of formation and the following algebraic relationship:
$$\Delta H = \sum \Delta H_f^{\circ}(\text{products}) - \sum \Delta H_f^{\circ}(\text{reactants})$$
17. Enthalpy diagrams based on reaction mechanisms can be used to gain understanding about the reasons why certain physical and chemical changes occur.

- Bond dissociation energies are useful in allowing one to predict the heat of reaction for specific chemical reactions.
- Optional:** When the atoms, molecules, or ions in a system become more disordered (more randomly distributed) during any change, the property known as entropy ( $S$ ) is said to increase. The change in entropy is symbolized as  $\Delta S$ . Changes will occur spontaneously when the difference between the enthalpy change ( $\Delta H$ ) and the absolute temperature ( $T$ ), multiplied by the entropy change ( $\Delta S$ ) gives a negative value. This difference is known as the free energy change and is symbolized by  $\Delta G$ . Thus,  $\Delta G = \Delta H - T\Delta S$ . Reactions occur spontaneously (proceed more in the forward direction) when  $\Delta G$  is negative; reactions are not spontaneous (proceed more in the reverse direction) when  $\Delta G$  is positive. When  $\Delta G = 0$ , the system is at equilibrium.

- Chemical equations (*Stoichiometry* module)
- Mole concept and stoichiometry (*The Mole* and *Stoichiometry* modules)
- Solution concentration and molarity (*Solutions* module)
- Kinetic molecular theory of gases (*Gases* module)
- Molecular shapes and structure (*Molecular Geometry* module)
- Difference between temperature and heat (*Gases* module)

## RELATED CONCEPTS

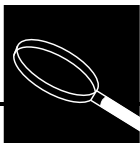
- Manipulative skills: Use of graduated cylinders, balances, thermometers.
- Mathematical and conceptual skills:
  - Problem solving techniques (*Introductory* and *Stoichiometry* modules)
  - Dimensional analysis and unit conversions (*Introductory* module)
  - Estimating answers
  - Simple algebra—combining several equations

## RELATED SKILLS

After completing their study of thermochemistry, students should be able to:

- differentiate between a system and its surroundings.
- determine the direction of heat flow in a system that undergoes a temperature change.
- differentiate among endothermic, exothermic, and adiabatic changes.
- define the joule, the calorie.
- differentiate among heat capacity, specific heat capacity, and molar heat capacity.
- given an initial temperature, final temperature, and mass of water involved in a water calorimetry experiment,
  - determine if heat was absorbed or evolved.
  - calculate the heat change in that system.
- write the heat term involved in a reaction including it as part of the chemical equation and using  $\Delta H$  notation.

## PERFORMANCE OBJECTIVES



8. discuss an exothermic (or endothermic) reaction in terms of heat changes and potential energy differences.
9. differentiate among gases, liquids, and solids in terms of the three types of kinetic energy—vibrational, rotational, and translational.
10. discuss the total energy of any system.
11. compare on a relative scale the energies involved in physical, chemical, and nuclear changes.
12. using a table of standard heats of formation, or known heats of reaction, use Hess's Law to calculate the heat of reaction for a reaction with unknown  $\Delta H$ .
13. given a table of standard heats of formation, calculate  $\Delta H$  for a given reaction.
14. determine the nature of a reaction, *i.e.*, exo- or endothermicity, from its enthalpy diagram.
15. construct an enthalpy diagram given a description of the energetics of the reaction.
16. predict the heat of reaction when bond dissociation energies are given.
17. **Optional:** Discuss entropy as it relates to a specific chemical or physical change, *e.g.*, a phase change from liquid to gas.
18. **Optional:** Given a table of standard heats of formation and standard entropy values, calculate  $\Delta G$  for a given reaction.



# Concept/Skills Development

## Activity 1: The Heat of Combustion of a Candle

### Introduction

A candle gives off light and heat energy as it burns. You might wonder, how much heat. To find out, you will use the definition of the unit of heat energy, the joule, and measure the temperature change produced in a known mass of water by a burning candle. You will also determine the mass (in g) of candle that has burned, so your results can be compared to those of your classmates in a meaningful way. Your final calculated result, in J/g, is the heat of combustion.

### Purpose

To determine the quantity of heat given off when a candle burns (combusts).

### Safety

1. Wear protective goggles throughout the laboratory activity.
2. A candle flame is hot, and the can holding the water (as well as the water) may also get hot, so exercise caution to avoid burning your hands.

### Procedure

To make optimal use of your time, you should collect the data in the order that follows:

1. Determine the mass of the candle to the nearest 0.01 g.
2. Measure the temperature ( $^{\circ}\text{C}$ ) of cold water provided by your teacher in your "calorimeter" (tin can).
3. Record these two measurements, then immediately light the candle and place it under the can of cold water (see Figure 1).
4. Record room temperature ( $^{\circ}\text{C}$ ).
5. Stir the water gently with a stirring rod, not the thermometer, and allow the candle flame to heat the water as far above room temperature as it was below room temperature at the start. (For example, if the water was at  $5^{\circ}\text{C}$  when heating began, and room temperature is  $25^{\circ}\text{C}$ , heat the water to a temperature of  $45^{\circ}\text{C}$  before blowing out the candle.) Keep stirring gently and record the highest temperature reached by the water as its final temperature.

## LABORATORY ACTIVITY: STUDENT VERSION

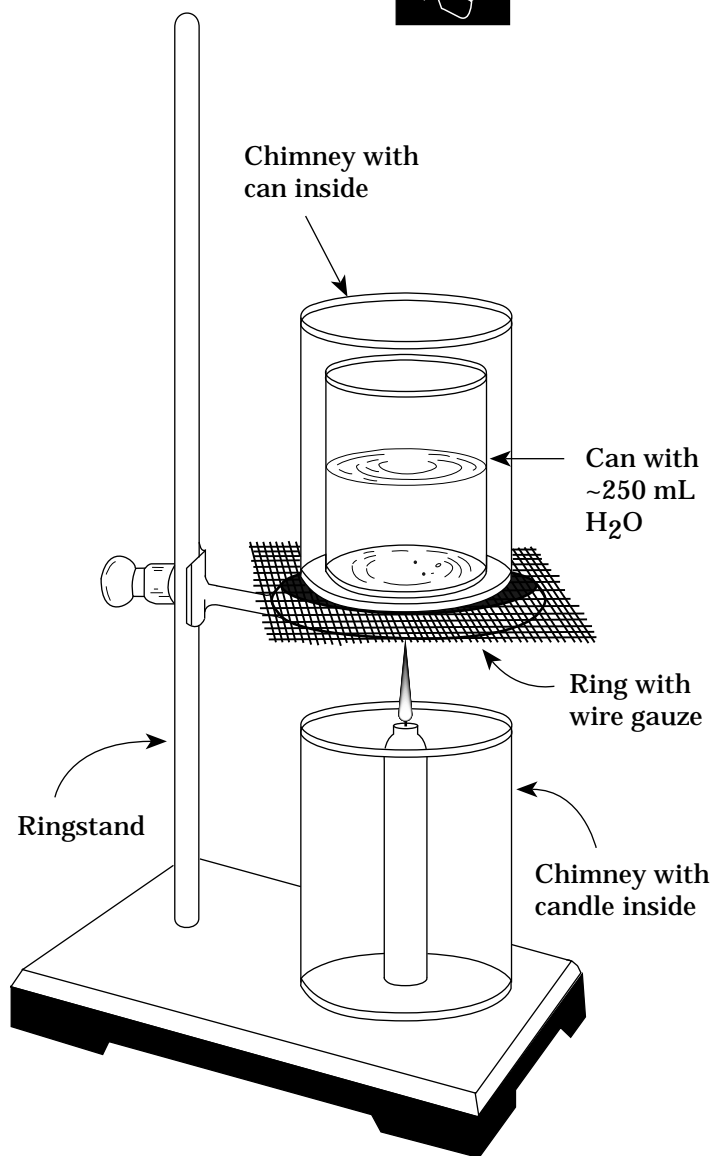


Figure 1. Apparatus for Activity 1.



- Determine mass of the can with its heated water, but only to the nearest gram.
- Discard the water into the sink, dry the can, and determine the mass of the empty (and dry) can, again only to the nearest gram.
- Finally, find the mass of the candle, to the nearest 0.01 g. Be sure to include any hardened drippings.
- Thoroughly wash your hands before leaving the laboratory.

### Data Analysis

- Organize your six measurements into a neat, concise data table, assigning each a capital letter (A,B,C,D,E,F).
- Make a table of calculated results, and indicate how each calculation was made, using letters. For example, if the mass of the candle before burning is represented by “A,” and its mass after burning by “B,” then the mass of candle burned, say “L” equals A – B. Place the numerical result of this calculation under its symbol. Be sure to include the uncertainties, both in the “Raw Data” (your measurements) and your calculated results.
- Think about the procedural directions you were given—the “strategy” of the activity. Why were you given cold water to heat instead of room temperature water? Why heat it up as far above room temperature as it initially was below that temperature?
- Now think about uncertainties, both in your original six measurements and the calculated results generated from them. How many places are justified in your values for the heat absorbed by the water, and for the heat of combustion of the candle?
- Finally, think about sources of error. How did your teacher attempt to minimize them? Will your experimental heat of combustion value due to these errors be too high or too low (relative to a standard heat of combustion)? Try to support your answer.

### Data Tables

#### Student Results

Raw Data	Candle ( $\pm 0.01$ g)		Tin can ( $\pm 1$ g)		Temp. ( $\pm 0.5$ °C)	
	Before	After	Empty	+ Water	Before	After
	A	B	C	D	E	F

#### Calculated Results

	Mass of candle burned (g)	Mass of water warmed (g)	Temp. change of water (°C)	Heat absorbed by water (J)	Heat of combustion of candle wax (J/g)
	$L = A - B$	$M = D - C$	$N = F - E$	$O = M \times N \times 4.184 \text{ J/g}^\circ\text{C}$	$P = \frac{O}{L}$
Value					
Uncertainty					
% Uncertainty					

### Implications and Applications

- Complete the following table of information for other fuels:

Fuel	Heat of combustion (kJ/g)	Molar heat of combustion (kJ/mol)	Temperature rise for 1 L water per 1 g fuel (°C/L/g)
C	32.8	394	7.84
CH <sub>4</sub>	55.7		13.3
C <sub>2</sub> H <sub>6</sub>		1560	12.4
C <sub>3</sub> H <sub>8</sub>	52.8		
C <sub>2</sub> H <sub>5</sub> OH		1390	
C <sub>6</sub> H <sub>12</sub> O <sub>6</sub>			3.70
H <sub>2</sub>		287	
U-235	8.00 x 10 <sup>7</sup>		

- Calculate the relative value of uranium as a fuel compared to coal (coal is mostly elemental carbon, C). (Use a ratio.)
- Discuss the advantages and disadvantages of using hydrogen gas as a fuel in cars.
- Discuss the advantages and disadvantages of using uranium as a fuel in nuclear reactors.

### References

Similar activities are presented in:

Merrill *et al.* (1987). *Experimental foundations lab manual* (pp. 39-40). Prentice-Hall.

American Chemical Society. (1993). *ChemCom: Chemistry in the community* (pp. 186-190). Kendall-Hunt.



**LABORATORY  
ACTIVITY:  
TEACHER  
NOTES**

***Activity 1: The Heat of Combustion of a Candle***

**Major Chemical Concept**

This laboratory exercise will allow students to conduct a calorimetry experiment, where they must decide what measurements they need to make, given the definition of a joule, and that:

$$\text{Heat (J)} = \text{Mass water (g)} \times C (4.18 \text{ J/g } ^\circ\text{C}) \times \text{Temperature change of water (} ^\circ\text{C)}$$

You should emphasize quantitative results and the accompanying uncertainties, given the very crude calorimeter used.

**Level**

All college preparatory classes of chemistry (and physics).

**Expected Student Background**

Students are expected to have been exposed to the concept of calorimetric measurement and uncertainty in measured data as well as in calculated results derived from these data.

**Time**

One or two class periods will be needed for this activity. The length of time available will determine how many different trials each student group will be assigned.

**Safety**

Read the *Safety Considerations* in the *Student Version*. Students should be cautioned about the hazards of an open candle flame and, eventually, hot water and hot tin can, at their laboratory bench. No other specific cautions are needed for this activity, except to follow the general safety guidelines already established for other activities.

**Materials** (For 24 students working in pairs)

**Nonconsumables**

- 12 Thermometers,  $-10$ - $110^\circ\text{C}$
- 12 Tin or aluminum cans, approximately 300 mL capacity
- 24 Larger cans, at least 1 L capacity (for "chimney")
- 12 Ring stands and rings
- 1 Laboratory balance, accurate to 0.01 g
- 1 Balance (0.1 or 1 g)

**Consumables**

- 12 Paraffin candles, 2-3 cm diameter x 18-20 cm length
- Matches
- Ice cubes for cooling water bath (*OPTIONAL*: Only used if teacher does not supply 1-2 L cold water.)

**Advance Preparation**

If possible, have 2 L cold water ready as a source to be poured into the smaller tin can for each group of two students. Each larger can (chimney) should have its top and bottom removed with a can opener, and several smaller cuts should be made at one end to allow air to enter when the larger can is placed over the candle apparatus in order to minimize wind disturbances, *etc.*

It is suggested that students also enter their measurements on a master sheet for later analysis and comparisons. Alternatively, a computer program or spreadsheet can be used to gather student data as well as the calculated heat of combustion. If a computer program or spreadsheet is to be used, have the computer ready for data entry by students.

### Pre-Laboratory Discussion

Encourage students to list essential measurements needed, and then talk about the order in which they should be made (see *Student Version*). Show the suggested set-up, including the use of the larger can to “protect” the candle and smaller can from room air currents. Point out the need to be able to calculate the mass of candle burned, giving the definition of heat of combustion (J/g). Stress the importance of collecting all drippings from the candle, and determining the mass of these along with the rest of the still unburned candle at the end of the activity. You may want to instruct students to place the candle in a small baking cup or on a tin can lid to contain the drippings, which, if used, must be part of the candle’s initial and final masses. Explain that it is essential that students quickly light the candle and place it so it begins to heat the can and cold water as soon as the temperature of the cold water is measured.

If a computer program or spreadsheet is to be used to collect data, explain how the data should be entered by students.

Encourage students to organize their data and calculated results so they know exactly how to use the data to obtain desired results. One way is to assign each measurement a letter and use these letters to indicate the calculations, as shown in the data tables.

Uncertainties need to be stressed so that students do not keep more significant figures in the calculated results than they should (see *Introductory* module).

Alternatively, you may choose to use the rule of significant figures (see *Introductory* module). If you choose this method, you will obtain a product (quotient) that has no expressed uncertainty; the implied uncertainty lies in the last significant figure that is allowed to be written. In the case of heat absorbed by water (O), 34,000 J, the calculated value (33,941 J) comes from the product of the specific heat of water,  $4.184 \text{ J K}^{-1}\text{g}^{-1}$ , mass (M), 208 g, and temperature change of water (N),  $39^\circ\text{C}$ . The best way to report this value is 34,000 J, because the value for N,  $39^\circ\text{C}$ , has only two significant figures. Thus according to the rule stated above, 33,941 J should be reported to two significant figures, or 34,000 J. (The uncertainty for 34,000 is located in the 1000s place, although the specific uncertainty is unknown for this process.) The proof is to work out the uncertainties, as above, and see that the last place in heat absorbed by water (O) and in the heat of combustion of candle wax (P) must be the thousands place, since that is itself uncertain. Students need to understand that to report the calculator answer of 33,941 J for heat absorbed by water (O) is a misrepresentation, since one cannot know it that precisely. To emphasize this, ask students how much the “missing” 59 J ( $34,000 \text{ J} - 33,941 \text{ J}$ ) would warm the more than 200 g water used. The approximate answer is only  $0.07^\circ\text{C}$ , obviously unreadable on their thermometers.

Point out that uncertainty comes into experimental strategy all the time. In this case, for example, ask students why they were told to find the mass of the candle before and after burning to the nearest 0.01 g, while finding the mass of the tin can both empty and full of water only to the nearest 1 g. Why read the tin can only to  $\pm 1 \text{ g}$  when, by taking a bit more time, these too could have been read to the nearest 0.01 g? The answer comes from uncertainties. Because loss of mass of the candle is small, even the  $\pm 0.02 \text{ g}$  uncertainty is a large relative uncertainty. Using expected student results, it is 2%, while the can masses measured to  $\pm 1 \text{ g}$  give a mass of water only 1% uncertain. Hence it makes no sense to spend time on those masses, since they already involve negligible uncertainties.



## Teacher-Student Interaction

Move around the laboratory, initially pouring approximately 200-250 mL of the cold water (but no ice!) into each of the smaller tin cans and checking the arrangement of the apparatus. Be sure students quickly weigh the candle ( $\pm 0.01$  g) and measure the water temperature. (The temperature initially should be approximately 4-5 °C.) Make sure students also immediately begin heating the cold water with the candle flame. As heating proceeds, remind students to stir gently, and to blow the candle flame out when the water temperature reaches approximately 45 °C. They must record the highest temperature reached by the water as its final temperature. Be sure students record all necessary data.

## Anticipated Student Results

Raw Data	Candle ( $\pm 0.01$ g)		Tin can ( $\pm 1$ g)		Temp. ( $\pm 0.5$ °C)	
	<i>Before</i>	<i>After</i>	<i>Empty</i>	<i>+ Water</i>	<i>Before</i>	<i>After</i>
	A	B	C	D	E	F
	45.97 g	45.06 g	61 g	269 g	6 °C	45 °C

Calculated results:

	Mass of candle burned (g)	Mass of water warmed (g)	Temp. change of water (°C)	Heat absorbed by water (J)	Heat of combustion of candle wax (J/g)
	$L = A - B$	$M = D - C$	$N = F - E$	$O = M \times N \times 4.184 \text{ J/g}^\circ\text{C}$	$P = \frac{O}{L}$
<b>Value</b>	0.91 g	208 g	39 °C	34,000 J	37,000 J
<b>Uncertainty</b>	$\pm 0.02$ g	$\pm 2$ g	$\pm 1$ °C	$\pm 1200$ J	$\pm 2100$ J
<b>*% Uncertainty</b>	$\pm 2$ %	$\pm 1$ %	$\pm 3$ %	$\pm 4$ %	$\pm 6$ %

\*More precisely, the % errors (E) are not additive but less since errors can be positive or negative:

$$E = \sqrt{E_1^2 + E_2^2}$$

## Answers to Data Analysis and Concept Development

1. See *Anticipated Student Results*.
2. See *Anticipated Student Results*.
3. As long as water is cooler than its room surroundings, heat will flow into water and warm it. When it is above room temperature, heat will leave the water and cool it. Since the temperature of the warmed water is as much above room temperature as the ice water was below room temperature, these errors cancel.
4. The uncertainties in the mass of candle burned ( $\pm 0.02$  g), mass of water warmed ( $\pm 1$  g) and temperature change are simply the sum of those in the data from which they are calculated. Uncertainties in heat absorbed by water and heat of combustion of candle wax are found from the sum of the percent uncertainties in mass of water warmed ( $\pm 1$  %) and temperature change ( $\pm 3$  %). Add the uncertainties (1% + 3%), and convert to uncertainty in heat absorbed by water by taking 4% of 34,000 J. (There are more precise computations for error analysis.)

5. All the heat given off by the candle is not measured. The can and surrounding air are also warmed by the candle, but not measured. This will result in an experimental heat of combustion that is too low. (Errors in measurement can result in random errors in the experimental heat of combustion.)

### Answers to Implications and Applications

1. Table of information for other fuels:

Fuel	Heat of combustion (kJ/g)	Molar heat of combustion (kJ/mol)	Temperature rise for 1 L water per 1 g fuel (°C/L/g)
C	32.8	394	7.84
CH <sub>4</sub>	55.7	891	13.3
C <sub>2</sub> H <sub>6</sub>	52.0	1560	12.4
C <sub>3</sub> H <sub>8</sub>	52.8	2330	12.6
C <sub>2</sub> H <sub>5</sub> OH	30.2	1390	7.22
C <sub>6</sub> H <sub>12</sub> O <sub>6</sub>	15.5	2790	3.70
H <sub>2</sub>	142	287	33.9
U-235	8.00 x 10 <sup>7</sup>	1.88 x 10 <sup>10</sup>	1.91 x 10 <sup>7</sup>

2. The relative value of uranium as a fuel compared to coal (mostly elemental carbon) is:

$$\frac{8.0 \times 10^7 \text{ kJ/g U-235}}{32.8 \text{ kJ/g C}} = 2.4 \times 10^6 \text{ times as much heat from U-235 as from an equal mass of coal.}$$

3. **Advantages:** Lightweight, compressible, plentiful (in seawater, if a source of energy to separate water into its elements is economical). **Disadvantages:** highly flammable (remember the Hindenberg!), “gas” tanks ruptured in collisions would result in cars becoming death traps
4. **Advantages:** >2 Million times as much energy available on a mass basis as can be obtained from burning coal, much energy from little mass, no air pollution as in burning coal, “filler-up” at the tanks would be very infrequent. **Disadvantages:** radioactivity hazards, long-term handling, and storage of nuclear waste is very unclear presently

### Post-Laboratory Discussion

Discuss the errors inherent in this activity. Get students involved in planning a strategy to minimize these errors. For example, ask them to justify starting with ice water that is to be warmed as far above room temperature as it was initially below that temperature. Why not start with room temperature and heat it to, say, 65 °C? (See Question 3 in *Answers to Data Analysis and Concept Development*.) An analogy may help here: Ask students what would happen in their house in winter if the furnace stops, or in summer if the air conditioner stops. (In the former, the heat would flow out of the house to the surroundings; in the latter, it would flow into the house from the outside.)

Ask if the heat measured (that entering the water) is all the heat given off by the candle. Point out that the can and surrounding air are also warmed by the candle, but not measured. Student values for the heat of combustion can therefore be expected to be too low. Indeed, the accepted value for paraffin wax is approximately 42,000 J/g.



Finally, be sure students understand the reason for dividing the number of joules released by the grams of candle burned. Point out that Student A might get 17,000 J by burning only 0.5 g candle, Student B might get 34,000 J by burning 1.0 g, Student C with 68,000 J from burning 2.0 g, yet each has the same heat of combustion: 34,000 J/g. This fact can be used to introduce  $\Delta H$  units of kJ/mol later on in the module.

### **Possible Extensions**

Ask students to design activities to measure heats of combustion of other fuels, such as sterno, rubbing alcohol and kerosene. *NOTE: For each of these materials, the mass loss during burning could be determined.* What about natural gas? How could results be standardized? The mass loss by the gas would be unavailable if it is taken from the laboratory gas jet, but the mass could be determined by collecting a sample over water. How about timing? Divide joules by seconds used for heating? Could these units be converted to joules per gram? How?

Ask students to research the procedure and apparatus used to obtain the “standard” values for heat of combustion located in handbooks, such as the *CRC Handbook of Chemistry and Physics*.

### **Assessing Laboratory Learning**

Give students raw data from another fuel’s heat of combustion experiment and ask them to design and complete a suitable “Calculated Results” table, ending with a heat of combustion as well as its uncertainty. Use data in *Answers to Implications and Applications* (Answer 1) to generate these “raw data.”

## Activity 2: Heats of Reaction

### Introduction

In this activity, you will observe and measure the heat of reaction for three different, but related, reactions. You will use your knowledge of the definition of heat to help you calculate the heat involved in each of these reactions. You will use the temperature change of a specified quantity of water contained in a glass Erlenmeyer flask, which will serve as the calorimeter. By knowing the mass of water and glass involved in the evolution or absorption of heat as well as the temperature change, you will be able to determine how much heat was exchanged in each of the reactions. These calculations will serve as a self-check on your experimental values.

To simplify the calculations somewhat, you will assume that the heat of reaction will raise the temperature of both the solution and the flask. We will neglect any heat lost to the surroundings. Recall that it takes 4.18 J to raise the temperature of 1 g water 1 °C. You may assume that the specific heat of solution is the same as that of pure water, 4.18 J/g °C. Glass has a specific heat of 0.84 J/g °C. The density of water is approximately 1.0 g/cm<sup>3</sup>; thus the volume indirectly gives you the mass of water involved. You will use this information to calculate the heat of reaction.

### Purpose

To determine the heat of reaction for several chemical reactions and then to use these values to calculate another  $\Delta H$ .

### Safety

1. Wear protective goggles throughout the laboratory activity.
2. The sodium hydroxide pellets used in this activity are very caustic. It is imperative that you do not touch these pellets, as serious chemical burns could result. If sodium hydroxide is spilled on the hands or other area of the body, flush with running water for at least 10 min, or until the soapy feeling is gone. Notify your teacher immediately. Sodium hydroxide splashed into the eye can cause serious damage, even after irritation has stopped. Flush eyes in the eye wash fountain for at least 15 min. Have someone notify your teacher immediately.
3. The solutions of hydrochloric acid involved can be irritating to skin, eyes, and mucous membranes. If you spill any on yourself, immediately flush the area with water for several minutes and notify your teacher. If any is splashed in your eyes, flush your eyes with water immediately and continue doing this for at least 15 min. Use the eye wash fountain if one is available in the laboratory.
4. The use of a mercury thermometer adds the potential danger of a mercury spill. Mercury is toxic and very difficult to clean up. If someone should break a mercury thermometer, notify your teacher immediately so that the mercury can be cleaned up completely and disposed of properly.
5. The solutions at the end of the activity can be safely poured down the drain, and flushed with tap water.

### Procedure

#### Part I

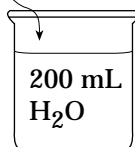
1. Determine the mass of a 250-mL beaker, to the nearest 0.01 g.
2. Pour 200 mL ( $\pm 1$  mL) of cool tap water into the beaker. Stir slowly and carefully with a thermometer until a constant temperature is reached (approximately room temperature). Record this temperature ( $\pm 0.2$  °C).

## LABORATORY ACTIVITY: STUDENT VERSION

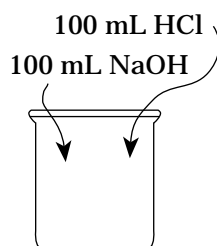
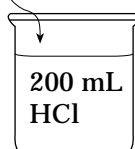




2gNaOH(s)



2gNaOH(s)



3. Determine the mass of about 2 g solid sodium hydroxide (NaOH), to the nearest 0.01 g. In order to do this as quickly as possible to prevent the solid from absorbing moisture from the air (and thus increasing its mass), your teacher will give you explicit instructions about how many pellets you will need to approximate 2 g (to stay within the range of 1.9 to 2.1 g).

4. Pour this measured sample of solid NaOH into the 200 mL of water already in the beaker. Stir the beaker vigorously (but avoid splashing) until the pellets are all dissolved. Measure the temperature and record the highest value observed.

5. Now dispose of this solution, and rinse the beaker in preparation for *Part II*.

### Part II

1. Repeat Steps 1, 2, and 3 from *Part I*, except in place of the tap water used in Step 2, substitute 200 mL of 0.25 M hydrochloric acid (HCl) solution.

3. Now repeat Steps 4 and 5 from *Part I*, adding the measured NaOH to the 0.25 M HCl.

### Part III

1. Measure 100 mL of 0.50 M HCl into a 250-mL beaker.

2. Measure 100 mL of 0.50 M NaOH into a second 250-mL beaker.

3. Measure the temperature of each of the solutions and record. (Rinse the thermometer before transferring it from one solution to the other.)

4. Pour the NaOH solution into the acid solution. Mix quickly and record the maximum temperature reached.

5. Thoroughly wash your hands before leaving the laboratory.

### Data Tables

	<b>Part I</b> <b>Reaction 1</b>	<b>Part II</b> <b>Reaction 2</b>
Mass of beaker, g		
Mass of weighing paper, g		
Mass of paper + NaOH, g		
Initial temperature, °C		
Final temperature, °C		

	<b>Part III</b> <b>Reaction 3</b>
Initial temperature HCl soln, °C	
Initial temperature NaOH soln, °C	
Final temperature mixture, °C	

## Data Analysis and Concept Development

### Part I

Fill in the results table below with the information from the following calculations.

1. Calculate the change in temperature,  $\Delta T$ , for the reaction.
2. Calculate the quantity of heat that is absorbed by the solution. (Remember that  $Q = m \times C \times \Delta T$ .) Assume that the density of all solutions is 1.00 g/mL.
3. Calculate the quantity of heat absorbed by the beaker.
4. Calculate the total quantity of heat absorbed by the system.
5. Calculate the amount of NaOH used (in moles).
6. Calculate the total quantity of heat involved per mole of NaOH.
7. Express your results as the heat of reaction,  $\Delta H$ , in kJ/mol.

### Part II

Fill in the results table for your data in *Part II*. The calculations are the same as for *Part I*.

### Part III

Fill in the results table for your data in *Part III*. The calculations are the same as for *Part I*.

### Results Table

	<i>Part I</i> Reaction 1	<i>Part II</i> Reaction 2	<i>Part III</i> Reaction 3
1. Change in temperature, °C			
2. Heat absorbed by solution, J (200 g x 4.18 J/g °C x $\Delta T$ )			
3. Heat absorbed by beaker, J (Mass beaker x 0.84 J/g °C x $\Delta T$ )			
4. Total heat absorbed, J (Heat soln + heat beaker)			
5. Moles NaOH used in reaction			
6. Total heat per mole NaOH, kJ/mol NaOH			
7. Heat of reaction, kJ/mol NaOH			

### Implications and Applications

1. Compare  $\Delta H_2$  with the sum of  $\Delta H_1$  and  $\Delta H_3$ .
2. Calculate the percent difference between  $\Delta H_2$  and the sum of  $\Delta H_1$  and  $\Delta H_3$ . (Assume  $\Delta H_2$  to be the correct value.)



3. Write the net ionic equations for Reactions 1, 2, and 3.
4. Each of the reactions above involves a specific type of heat of reaction. Reaction 1, for example, represents the heat of solution of solid NaOH. Write similar statements for the other two reactions.
5. Suppose you had used 4.0 g NaOH in Reaction 1, instead of 2.0 g.
  - a. Calculate the quantity of heat involved.
  - b. How would this result affect the  $\Delta H$ , measured in kJ/mol, for this reaction?
6. How could we make use of the concept of heats of reaction in our everyday lives?

### References

Similar activities are presented in *Chemistry: Experimental Foundations Lab Manual* (Merrill *et al.*, 1987, pp. 60-61) and *CHEMStudy Lab Manual* (1965, pp. 39-40).

**Activity 2: Heats of Reaction****Major Chemical Concept**

This laboratory activity will provide data for students to calculate heats of reaction for three different reactions, and then to compare them in order to demonstrate Hess's Law.

**Level**

All college preparatory classes in chemistry.

**Expected Student Background**

Students should have already performed calorimetry calculations involving mass of water heated and its temperature change. The goal of this exercise is to take calorimetry one step farther and to compare different  $\Delta H$  calculations to each other.

**Time**

All three parts of this activity can be completed in a 50-min period. Alternatively, specific parts could be assigned to individual laboratory groups, and their results could be combined to get a complete set of class data.

**Safety**

Read the *Safety Considerations* in the *Student Version*. You should be vigilant in observing student behavior throughout this series of activities. Be sure students have been shown the location of the eye wash facility prior to the start of the activity.

**Materials** (For 24 students working in pairs)**Nonconsumables** (per laboratory team)

24 Beakers, 250-mL  
12 Thermometers  
12 Graduated cylinders, 100-mL  
Centigram balance  
12 Scoopulas  
12 Stirring rods

**Consumables**

*NOTE: Since this is a quantitative experiment, the solutions must be made up carefully to the proper concentrations. Ideally, the class solutions should be standardized.*

Sodium hydroxide pellets, NaOH, 48 g  
0.25 M Hydrochloric acid, HCl, 2 L (42 mL conc HCl diluted to 2 L)  
0.5 M Hydrochloric acid, HCl, 1 L (42 mL conc HCl diluted to 1 L)  
0.25 M Sodium hydroxide, NaOH, 1 L (10 g NaOH per 1 L solution)  
0.5 M Sodium hydroxide, NaOH, 1 L (20 g NaOH per 1 L solution)  
Tap water, 200 mL  
Weighing paper

**Advance Preparation**

1. Solutions of HCl and NaOH must be prepared at least one day in advance so that they will be at room temperature on the day of the experiment. Both dilution reactions release heat, so the solutions will become hot when first prepared. Time must be allowed (at least overnight) for this heat to dissipate.

**LABORATORY  
ACTIVITY:  
TEACHER  
NOTES**



2. Since rather large volumes of acid and base solutions will be needed, you may want to employ siphons in large 4-L jugs with rubber tubing and pinch clamps to dispense the liquids to students.
3. The mass of the moisture picked up by the NaOH is apparently not significant, even though it appears that the pellets become quite wet with time. Results should not be affected by the slight (<1%) mass gain.
4. Ask students all to use the same labeled graduated cylinder (or two) for one solution, and keep it (them) at each jug, to keep dilution to a minimum and ensure that students are not contaminating their solutions *via* unwashed graduated cylinders. This method will require students to have only one cylinder at their laboratory station, to measure the water for *Part I*.

Alternatively, you may want to have another siphon in a jug of slightly cool water that you have cooled ahead of time as a fourth solution. Assigning a labeled graduated cylinder or two at this fourth jug will negate the need for any cylinders at individual laboratory stations. Using this method, rinse each graduated cylinder first with a small volume of the assigned liquid and discard the liquid. From that point on, the only liquid that should be in that cylinder will be the correct (assigned) one.

### Pre-Laboratory Discussion

1. Avoid mentioning Hess's Law to allow students to arrive at this "discovery" themselves.
2. Tell students the number of pellets required to approximate 2.0 g. Explain the method they are to use to clean up spilled pellets.
3. Warn students of the dangers of both hydrochloric acid and sodium hydroxide solutions, as well as solid sodium hydroxide.
4. Discuss with students the appropriate time for drying glassware (only when its mass needs to be determined). This means that the flask only needs to be dried once—the first time it is cleaned, because its mass will be determined immediately thereafter. Drying only when necessary will save students time.
5. Discuss with students the method you are using to dispense the solutions.

### Teacher-Student Interaction

If assigned pre-laboratory preparations include making a data table, move around the room to ensure students have done this. (This is one sign that advanced preparations have been done by the student.) Also check to see that students understand the directions, especially any given verbally that deviate from the printed version.

### Anticipated Student Results

	<b><i>Part I</i></b> <b>Reaction 1</b>	<b><i>Part II</i></b> <b>Reaction 2</b>
Mass of beaker, g	125 g	125 g
Mass of weighing paper, g	0.89 g	0.88 g
Mass of paper + NaOH, g	2.97 g	2.80 g
Initial temperature, °C	24.0 °C	23.8 °C
Final temperature, °C	26.4 °C	29.6 °C

*Remarks*

This demonstration is similar to the one above involving two solids, except that this is merely a solvation processes. As the endothermic solvation process proceeds, enough heat is lost by the water on the cork to freeze the water. The ice fuses the watch glass and cork together.

**Demonstration 3: Hot and Cold Packs***Purpose*

To illustrate that dissolution of solids can cause heat changes that are easily detected by feeling the container.

*Materials*

Calcium chloride, anhydrous,  $\text{CaCl}_2$ , 30 g

Ammonium chloride,  $\text{NH}_4\text{Cl}$ , 50 g

Water

2 Zip-closure bags

2 Vials with caps, 30-mL

*Safety*

The hot pack gets hot enough to cause discomfort. If you pass the cold and hot packs around the class, caution students not to hold the hot pack solution in their hand for too long. Contents can be flushed down the drain after the demonstration.

*Procedure*

Hot pack:

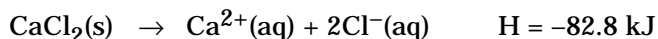
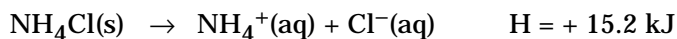
1. Measure approximately 30 g  $\text{CaCl}_2$  into a sandwich-sized zip-closure bag.
2. Fill a small vial with about 25 mL of water and cap it.
3. Insert the sealed vial into the bag and seal the bag.
4. Keeping the zip-closure bag sealed, pop the cap off the vial from the outside. Pour the water into the solid.
5. Shake the bag to mix the contents.
6. Pass the system around the class and have students observe the temperature change.

Cold pack:

1. Repeat Steps 1-6 above using 50 g  $\text{NH}_4\text{Cl}$  (instead of  $\text{CaCl}_2$ ).

*Remarks*

The heats of reaction for the two reactions above are:



Relate these two reactions to the real-world use of hot and cold packs, *e.g.*, in sports injuries, hunting, *etc.*

*Reference*

Marsella, G. (1987, February). Hot and cold packs. *ChemMatters*, 5(1), 7-12.



## Demonstration 4: Gallon of Gas Reaction (Methane Gas Can Explosion)

### Purpose

To safely demonstrate a potentially explosive reaction involving the rapid combustion of natural gas.

### Materials

New, empty 1-gallon paint can  
Masking tape, two 2" pieces  
Rubber tubing  
Methane or natural gas (mostly methane—propane won't work)  
Matches  
Long wax taper  
Lucite explosion shield

### Safety

Position the can so that the propelled lid will not impact with any student or breakable object (light fixtures, *etc.*). Use the explosion shield for protection (as described).

### Procedure

1. Drill a 3/16" hole in the top of the lid of the paint can.
2. Drill a 1/2" hole about 1" above the bottom of the can.
3. Place the lid on the can and press it down tightly.

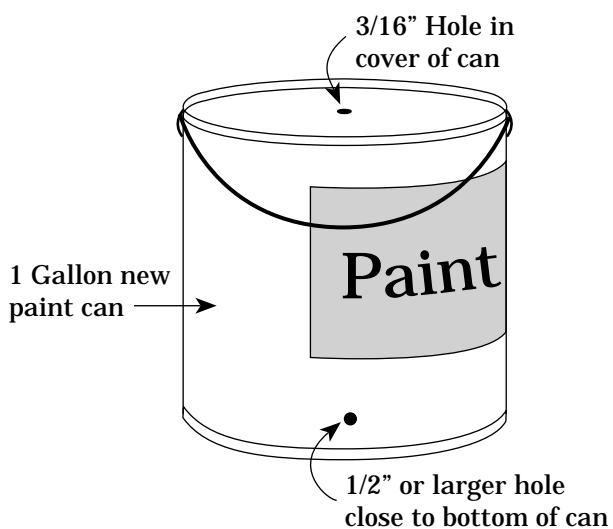


Figure 2. Diagram for Demonstration 4.

4. Attach one end of the rubber tubing to your laboratory gas jet or methane supply and the other end into the can (either hole). Since methane is lighter than air, the can should be above the point of inlet for methane.
5. Fill the can with gas and tape both holes shut with tape.
6. Place the can in an area that is not exposed to drafts, put the explosion shield in front of the can, and after removing both pieces of tape, light the gas at the top hole with the long taper.
7. Observe the size and color of the flame at the beginning, and compare with size and color as the reaction progresses. *NOTE: As the demonstration progresses, the flame becomes smaller and less luminous, even blue in color. Eventually it is just at the surface of the opening in the top of the can. At this point, any air current or movement of the can will cause the flame to extinguish. To prevent this, the can should be kept away from students and demonstrator alike, and surrounded, if possible, by a transparent plastic explosion shield to keep drafts of air away from the system.*

8. When the reaction is over, the can can be replenished immediately with gas in preparation for your next class. You can prepare the can in advance (even overnight), but be sure to tape the holes shut tightly.

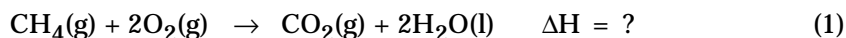
**Remarks**

Even as the reaction proceeds, you can discuss the heat of reaction involved. This is a great example to illustrate Hess's Law of Heat Summation—students can see the results!

The reaction time is somewhat reproducible, typically about 10-12 min. If you time the demonstration as you practice, you will be able to predict rather closely the time of "extinguishing."

Equations involved are:

Overall reaction:

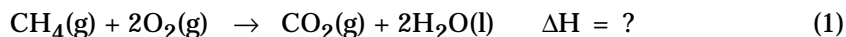


Reactions for calculating the heat of reaction:



Reverse Reaction 1 above, and add the three reactions to obtain the overall equation. Then add  $\Delta H$ 's to get  $-890.35 \text{ kJ}$  for the overall reaction.

Answer:



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

$$= [\Delta H_{\text{f}}^{\circ} \text{CO}_2 + 2\Delta H_{\text{f}}^{\circ} \text{H}_2\text{O}] - [\Delta H_{\text{f}}^{\circ} \text{CH}_4 + 2\Delta H_{\text{f}}^{\circ} \text{O}_2]$$

$$= \left[ 1 \text{ mol CO}_2 \times \frac{-393.51 \text{ kJ}}{1 \text{ mole CO}_2} + 2 \text{ mol H}_2\text{O} \times \frac{-285.83 \text{ kJ}}{1 \text{ mol H}_2\text{O}} \right] -$$

$$\left[ 1 \text{ mol CH}_4 \times \frac{-74.81 \text{ kJ}}{1 \text{ mol CH}_4} + 2 \text{ mol O}_2 \times \frac{0 \text{ kJ}}{1 \text{ mol O}_2} \right]$$

$$= (-965.17 \text{ kJ}) - (-74.81 \text{ kJ})$$

$$= -890.36 \text{ kJ}$$

Discussion can involve much more detail, including an enthalpy diagram (see *Transparency Masters 6 and 7* in the *Appendix*), and calculation of heat of reaction on the basis of bond dissociation energies.

## Demonstration 5: Methanol—Tennis Ball Cannon

**Purpose**

To demonstrate the explosive nature of the reaction of methanol vapors, and to relate this to enthalpy and entropy changes.

**Materials**

Several cans of the same diameter, such as tennis ball cans or food cans of small diameter

1 Tennis ball (or Nerf ball)

Electrician's tape

Dropping bottle of methanol

Matches, long (if long matches are not available, use a wax taper)

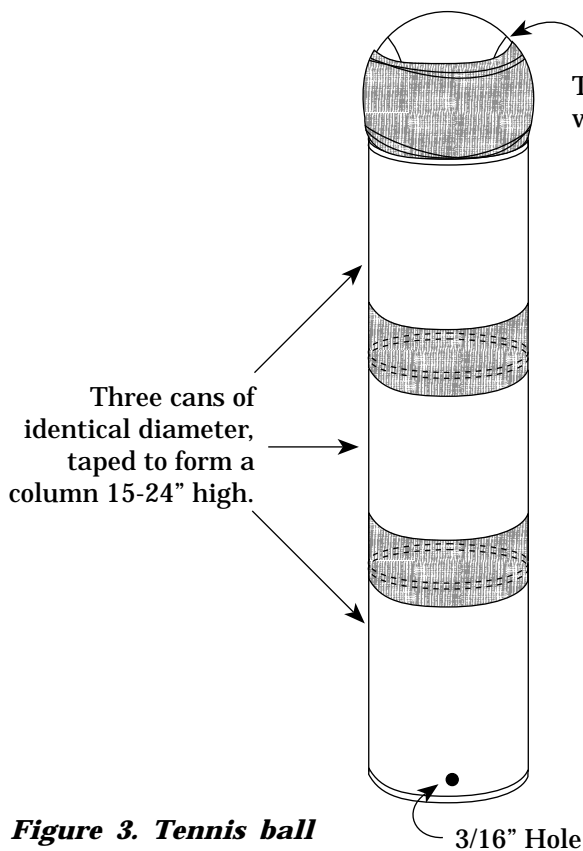
Ringstand and support clamp



### Safety

Be very careful to aim the “cannon” away from any person in the room. An explosion shield should be used in front of the taped stack of cans to prevent a rupture from propelling a can into the audience. Do not use plastic “cans”!

### Procedure



Tennis ball wrapped with electrical tape.

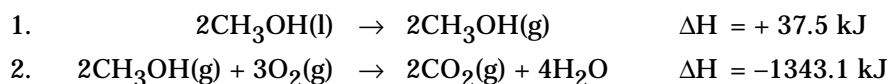
**Figure 3. Tennis ball can diagram.**

1. Cut both the top and bottom from two small diameter cans, and one metal tennis ball can or two tennis ball cans.
2. Cut only one end off one can. (One end should be closed.)
3. Tape the (two or) three cans together with electrician's tape, going around each can several times to ensure a tight seal. Several cans of any height with a diameter approximately the same as that of a tennis ball can be taped together. (Apparently tennis ball cans are no longer made of metal, and the plastic varieties don't work very well.)
4. Drill a 5/16" hole about 1/2" from the bottom of the can with the one end remaining. Make sure the inside of the upper open can is filed smooth. (File the inside if necessary.)
5. Wrap the tennis ball with electrical tape until the ball fits snugly into the open end of the taped set of cans (the “cannon”).
6. Place the tennis ball snugly into the opening of the stack of taped cans.
7. Squeeze two eyedroppers of methanol into the hole at the base of the bottom can.

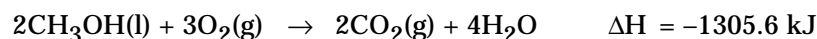
8. Overturn the can and allow the methanol to evaporate. While you await the evaporation, mount the can with a clamp to a ringstand on the table. (You might want to tilt the cannon at an angle so that the ball's trajectory will not intercept anyone or any fragile object.)
9. Light a long match or wax taper and hold it at the hole in the bottom can. The methanol-air mixture will explode and drive the ball into the air. The effect is greatest when the tape-wrapped ball is tightly fitted into the top can.
10. An immediate repetition of the demonstration usually is a dud, illustrating that a new supply of air (oxygen) is necessary for the reaction to recur.

### Remarks

At 298.2 K (25 °C):



Overall,



The large increase in gas volume (resulting in the increased pressure that propels the ball) is correlated to an overall entropy change of +202.9 J/K.

## **Demonstration 6: Heat, Temperature, and Heat Capacity In Balloons**

### *Purpose*

To illustrate the large heat capacity of water as compared to air.

### *Materials*

Balloons, two large, dark (black or violet)  
Water  
Candle  
Two ringstands

### *Safety*

Don't leave the flame under the second balloon for too long, or the balloon may develop a weak spot and burst. Discard balloons after use.

### *Procedure*

1. Blow up one balloon to a diameter of 6-8 inches and tie a knot or seal it off by some other method.
2. Fill the second balloon part way with some water (approximately 50-100 mL will do) before blowing it up to the same approximate diameter as the first balloon. Tie this one off in the same way you did the first one.
3. Suspend the two balloons almost at the top of two separate ringstands. (If you have done all this correctly, unknowing students should not be able to detect the pool of water in the bottom of the second balloon.)
4. Ask students what they think will happen (prediction) when you hold a lighted candle under one balloon. Light a candle and test their hypothesis, holding the lit candle under the balloon filled only with air.
5. Now discuss (ostensibly) reproducibility of results and suggest the need to repeat the activity with the second balloon. Hold the lighted candle at the bottom of the second balloon where it will touch the center of the pool of water. Students will expect this balloon to burst also, but it will not.

### *Remarks*

The second balloon will not burst because of the high heat capacity of the water. High heat capacity means that the water will absorb large quantities of heat (from the candle flame) without changing its temperature appreciably. The temperature of the burning candle is high (high enough to burst the material of the first balloon), whereas the temperature of the water in the second balloon is much lower (probably 20-25 °C). The temperature of the flame could be measured with a metal probe thermometer, while the water temperature can be easily measured by a standard laboratory thermometer.

Heat is that form of energy that flows across the boundary (synthetic rubber balloon) from the external high temperature to the lower temperature of the water. The water, rather than the balloon, absorbs the heat.

A similar activity can be done using a paper cup as the container in which to boil water. If one maintains a low flame, even on a burner, the water will boil with little or no charring of the paper. Be sure you use a cup that is not coated with wax or plastic to avoid the wax or plastic burning first!



## Demonstration 7: Burning Candles—Reaction Mechanisms and Concurrent Reactions—Oxidation

### *Purpose*

To develop, on the submicroscopic level, an understanding of what is occurring when a candle burns.

### *Materials*

Several candles  
One candle with exposed wick at both ends  
Florence flask (or round bottom flask) containing ice  
Wide-mouth Erlenmeyer flask  
Two support clamps  
Matches  
Ice

### *Safety*

As always, be careful when using flammable substances. Be especially cautious when burning your candle at both ends!

### *Procedure*

#### **Part 1**

1. Light a candle and allow it to form a good-sized pool of wax.
2. Light another match, blow out the candle and then hold the lighted match above the wick, not touching the wick. The candle should relight without touching the match.

#### **Part 2**

1. Hold the round flask containing the ice a few inches above a lighted candle. A black deposit of soot (unburned carbon) will start forming, and perhaps moisture will condense.

#### **Part 3**

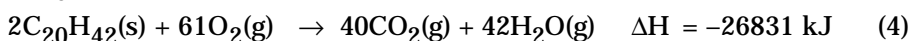
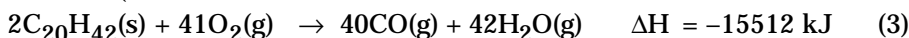
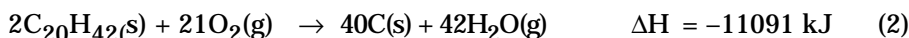
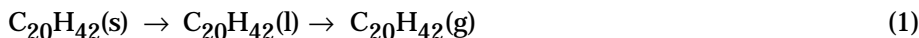
1. Place a clamp on the neck of the Erlenmeyer flask and hold the upended flask so that the neck opening is above the burning candle. Moisture will collect on the **inside** walls of the flask.

#### **Part 4**

1. Light one end of the double-wick candle holding it with a clamp, while concealing the other wick from the class.
2. Ask them, “How could you measure the rate at which a candle burns?”  
[*Measure height change, measure mass decrease, etc.*]
3. Then ask them, “How would you increase the rate at which the candle burns?” [Use pure oxygen gas.]
4. Then light the candle at the other end to show yet another way to increase the rate. A quote from Dr. Hubert N. Alyea, Professor Emeritus of Princeton University, fits here: “If you burn your candle at both ends, you will only live half as long . . . but you’ll have more than twice as much fun!”

## Remarks

The concurrent thermochemical reactions are:



*Part 1* of the procedure can be explained by observing Equations 1, 2, and 3 above, all of which contain unburned substances [ $\text{C}_{20}\text{H}_{42}(\text{g})$ ,  $\text{C}(\text{s})$ , and  $\text{CO}(\text{g})$ ] that are available for further burning if the energy of activation is provided (as by the match).

*Part 2* of the demonstration shows Reaction 2 above, where unburned carbon deposits on the flask bottom.

*Part 3* of the demonstration shows the production of water vapor that condenses inside the flask. This is shown in Reactions 2, 3, and 4.

*Part 4* of the demonstration involves the addition of oxygen to the reaction system. Oxygen is the limiting reagent in the burning of a candle, as Reactions 2 and 3 show. Only when an abundance of oxygen is present does the paraffin burn completely to carbon dioxide and water vapor. Typically, Reaction 2 above is the major combustion reaction for candles, as evidenced by the yellow flame (incandescent carbon particles) and the evolution of smoke or soot (unburned carbon).

Questions you might want to ask students:

1. What are the functions of the wick? [*To light the candle, to melt the wax, to draw up molten wax by capillary action, to vaporize the wax.*]
2. How do we know that wax vapor is undergoing combustion? [*Remember the match above the just-burned wick?*]
3. Which is the major combustion reaction in the candle? [*Reaction 2*]
4. Why is this the major combustion reaction? [*Oxygen gas is the limiting reagent.*]
5. Discuss the mechanism of burning of the candle up to Reaction 2 above. [*paraffin(s)  $\rightarrow$  paraffin(l)  $\rightarrow$  paraffin(g)*]
6. What makes the mechanisms of Reactions 2, 3, and 4 difficult to determine? [*The sheer number of molecules involved in the reaction coupled with the complexity of the bond breaking and bond forming steps on a molecular level.*]
7. Which are consecutive phase changes? [*The two reactions included in Reaction 1 illustrate consecutive reactions.*]
8. Which are concurrent reactions? [*Reactions 2, 3, and 4 illustrate concurrent reactions.*]

Two questions that could be cross-referenced in *Links and Connections* section:

1. Contrast Reactions 2, 3, and 4 in terms of energy production and air pollution. [*Obviously, Reaction 4 produces the most energy (26831 kJ vs. 11091 kJ and 15512 kJ), and it is also the least polluting of the three reactions, although carbon dioxide is now known to contribute to the "greenhouse effect." The other two reactions each produce less energy per mole of paraffin burned, and they contaminate the air by producing dirty soot and carbon monoxide.*]



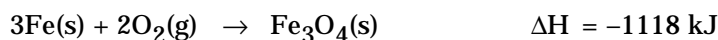
2. In the burning of natural gas (consisting of  $\text{CH}_4$ ,  $\text{C}_2\text{H}_6$ ,  $\text{C}_3\text{H}_8$ , and  $\text{C}_4\text{H}_{10}$ ) in a burner, gas stove, gas oven or gas heater, the blue flame indicates a much more energy-efficient burning process. Why is the efficiency of this process so much greater than that of the combustion of candle wax? Use the equations for the combustion reaction for paraffin to help explain this. Think mechanisms. [The main reason is the premixing of natural gas with air (oxygen). Gasoline (e.g.,  $\text{C}_8\text{H}_{18}$ ) burns more efficiently because it is premixed in the carburetor. The four gases involved in natural gas contain far fewer carbon and hydrogen atoms than paraffin, and thus occur in much less complex reaction mechanisms than paraffin. Also, since they are already gases, whereas paraffin is solid, far fewer complicated mechanisms are needed for bond breaking and reforming with natural gas than with paraffin.]

#### References

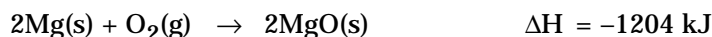
See *Demonstration 1* in the *Reaction Rates* module for other burning reactions that have thermochemistry applications.

$\Delta H$ 's for several of those reactions follow here:

Burning of steel wool:



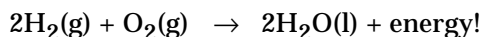
Burning of magnesium powder:



These reactions with additional data could be used to illustrate the Born-Haber cycle. Since gases are consumed in both reactions, the entropy changes are negative and are unfavorable driving forces for the reaction. The latter reaction is employed in flash bulbs and cubes. Boy Scouts are taught to light fires with the steel wool/iron reaction. The lycopodium powder reaction has been used by magicians under the name, "dragon's breath." These reactions can be also used to introduce balancing of oxidation-reduction reactions.

### Other Demonstration Ideas

#### 1. Explosive Reaction



See: Summerlin and Ealy. (1988). *Chemical demonstrations* (p. 27). American Chemical Society.

#### 2. Electrolysis of Water in Hoffman Apparatus

Do this demonstration in concert with the previous demonstration above, emphasizing the large quantity of electrical energy needed to produce (and thus stored in) the small amounts of hydrogen and oxygen gases involved in the demonstration (see *Demonstration 4* in *The Mole* module.) You might also want to explode bubbles of a mixture of hydrogen and oxygen to show that large quantities of energy are released when these two elements recombine to form water.

Hubert Alyea, Professor Emeritus in Chemistry from Princeton, a master of demonstrations, did a demonstration involving a hydrogen/oxygen "pop" bottle. The demonstration consisted of a stoichiometric ratio (2:1) of hydrogen and oxygen in a stoppered soda "pop" bottle. He lit the mixture (aim away from the audience), then explained that this is the reason it is called a "pop" bottle! He also showed that water was produced in the reaction by pouring

out several milliliters of water after the reaction. (He had poured the water in behind the scenes before the demonstration!) Much discussion has ensued recently concerning the safety of the demonstration and the probability of the bottle's exploding. No conclusion has been reached, but you would be wise to tape the entire bottle in clear tape or netting to minimize the risk. A plastic pop bottle might be a wise choice.

### 3. Spontaneous Exothermic Reactions

See *Chemical Demonstrations 41* "Exothermic Reaction: Calcium Chloride" (Summerlin & Ealy, 1988, p. 67) and *Chemical Demonstration 42* "Exothermic Reaction: Sodium Sulfite and Bleach" (Summerlin & Ealy, 1988, p. 68).

### 4. Entropy Increase during Endothermic Reaction

See Demonstration 26 in *Chemistry: Experimental Foundations, Teachers' Guide* (Parry et al., 1987, pp. 381-382).

## Key Questions

1. What is the meaning of the terms "surroundings" and "system"? [*The system is the portion of the universe under investigation. The surroundings are the rest of the universe.*]
2. What is the difference between heat and work? [*Heat is a form of energy that crosses a boundary whenever there is a difference in temperature at the boundary. The rate at which heat energy is transferred depends upon many factors including the thermal conductivity of the barrier and the temperature differential,  $\Delta T$ . Work is any quantity of energy that flows across the boundary of a system during a change in state and is completely convertible into the lifting of a mass in the surroundings.  $w = mgh$  (where  $w$  = work,  $m$  = mass,  $g$  = acceleration due to gravity, and  $h$  = height the mass was lifted); if the mass is lifted,  $h$  is positive and we say work has flowed into the surroundings.]*]
3. What is the difference between temperature, heat, and heat capacity? [*See Misconception 1.*]
4. What is the difference between an endothermic reaction and an exothermic reaction? Give at least one example of each. [*In endothermic reactions heat is added to the system, whereas in exothermic reactions the system loses thermal energy. Examples of these two types of reactions are given throughout this module.*]
5. Consider the bonds, C–C, C=C, and C≡C. Which bond requires the greatest amount of energy to break? Which bond is the shortest? [*The triple bond, C≡C, is the strongest because, in essence, there are three bonds to be broken here. For the same reason, the triple bond is the shortest bond, since the forces of attraction between these two carbon atoms is the strongest.*]

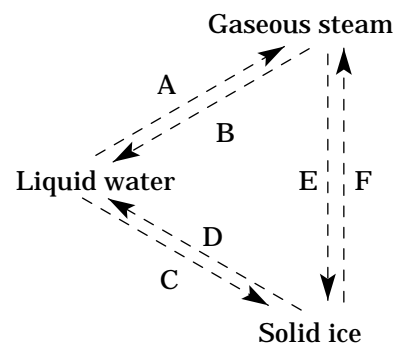
## GROUP AND DISCUSSION ACTIVITIES

<b>Data</b>	<b>Bond dissociation energy</b>	<b>Bond length</b>
C–C	348 kJ/mol	0.154 nm
C=C	614 kJ/mol	0.133 nm
C≡C	812 kJ/mol	0.120 nm



6. Examine the diagram:

- a. What names are given to the changes A, B, C, D, E, and F? [A = evaporation, B = condensation, C = freezing, D = fusion (melting), E = deposition (desublimation?), F = sublimation]



- b. Which changes are exothermic? [B, C, E]  
c. Which changes are endothermic? [A, D, F]  
d. Define the terms molar heat of fusion and molar heat of vaporization. [Molar heat of fusion is the heat required to convert one mole of solid to liquid. Molar heat of vaporization is the heat required to convert one mole of liquid to gas.]

7. Which would have the higher entropy?

- a.  $\text{H}_2\text{O}(\text{l})$  or  $\text{H}_2\text{O}(\text{g})$ ? [ $\text{H}_2\text{O}(\text{g})$ ]  
b.  $\text{Cl}_2(\text{g})$  or  $2\text{Cl}(\text{g})$ ? [ $2\text{Cl}(\text{g})$ ]  
c.  $\text{CO}_2(\text{s})$  or  $\text{CO}_2(\text{g})$ ? [ $\text{CO}_2(\text{g})$ ]  
d.  $\text{Ag}^+(\text{g})$  or  $\text{Ag}^+(\text{aq})$ ? [ $\text{Ag}^+(\text{g})$ ]

8. Which would have the higher energy?

- a.  $\text{H}_2\text{O}(\text{l})$  or  $\text{H}_2\text{O}(\text{g})$ ? [ $\text{H}_2\text{O}(\text{g})$ ]  
b.  $\text{Cl}_2(\text{g})$  or  $2\text{Cl}(\text{g})$ ? [ $2\text{Cl}(\text{g})$ ]  
c.  $\text{CO}_2(\text{s})$  or  $\text{CO}_2(\text{g})$ ? [ $\text{CO}_2(\text{g})$ ]  
d.  $\text{Ag}^+(\text{g})$  or  $\text{Ag}^+(\text{aq})$ ? [ $\text{Ag}^+(\text{g})$ ]

9. Write an equation for the complete combustion of:

- a. **Methane**  $\text{CH}_4(\text{g})$  [ $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$ ]  
b. **Propane**  $\text{C}_3\text{H}_8(\text{g})$  [ $\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$ ]  
c. **Butane**  $\text{C}_4\text{H}_{10}(\text{g})$  [ $2\text{C}_4\text{H}_{10}(\text{g}) + 13\text{O}_2(\text{g}) \rightarrow 8\text{CO}_2(\text{g}) + 10\text{H}_2\text{O}(\text{g})$ ]  
d. **Octane**  $\text{C}_8\text{H}_{18}(\text{l})$  [ $2\text{C}_8\text{H}_{18}(\text{l}) + 25\text{O}_2(\text{g}) \rightarrow 16\text{CO}_2(\text{g}) + 18\text{H}_2\text{O}(\text{g})$ ]

10. It takes 571.5 kJ of energy to decompose two moles of water. How much energy is needed to decompose 1000 g water?

$$1000 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \times \frac{571.5 \text{ kJ}}{2 \text{ mol H}_2\text{O}} = 15875 \text{ kJ} \cong 2 \times 10^4 \text{ kJ}$$

### Counterintuitive Examples

1. A spontaneous endothermic reaction seems to be in opposition to the tendency to minimum energy. It can be explained using the tendency toward maximum entropy. If the quantity,  $T\Delta S$  is positive and greater in magnitude than  $\Delta H$ , the reaction can be spontaneous ( $\Delta G < 0$ ) despite its endothermicity. ( $\Delta G = \Delta H - T\Delta S$ )
2. Household dehumidifiers do not cool a home, yet a dehumidified home feels cooler than a nondehumidified home, even though both homes are at the same temperature because evaporation of water from the skin carrying away heat from the body occurs more rapidly in a low humidity environment.

## Metaphors and Analogies

1. Use the analogy of a book on the floor vs. the same book on an upper shelf of a bookcase to demonstrate the energy of a substance changing as it reacts or changes.
2. See Item 1 in *Common Student Misconceptions*. Refer to the mechanical analogy for endothermic and exothermic reactions detailed in *Metaphors and Analogies* in the *Rates of Reaction* module.

## Language of Chemistry

**adiabatic process** process in which no heat exchange occurs between the system and its surroundings.

**bomb calorimeter** device used to measure the heat of a combustion reaction. The quantity measured is the heat of reaction at constant volume,  $q_v = \Delta E$ .

**bond energy** quantity of energy required to break a particular bond in one mole of gaseous molecules.

**calorimeter** device used to measure the quantity of heat exchanged between a system and its surroundings.

**calorimetry** measurement of heat changes.

**chemical energy** energy stored within the structural units of chemical substances.

**closed system** system that allows the exchange of energy (usually in the form of heat) but not mass with its surroundings.

**endothermic process** process that absorbs heat from its surroundings

**energy** capacity to do work or to produce change.

**enthalpy (H)** thermodynamic quantity used to describe heat content of a system at constant pressure.

**enthalpy change** difference in enthalpy between two states of a system at constant pressure.

**enthalpy of formation** enthalpy change accompanying the formation of a mole of a compound from the most stable forms of its elements.

**enthalpy of reaction** difference between the enthalpies of the products and the enthalpies of the reactants.

**enthalpy of solution** heat generated or absorbed when a certain amount of solute is dissolved in a certain amount of solvent.

**entropy (S)** direct measure of the randomness or disorder of a system.

**entropy change** extent to which the degree of order changes as the result of some process (Positive  $\Delta S$  means an increase in disorder of the system).

**entropy change of universe** total entropy change, both system and surroundings, for a given process. For every spontaneous change,  $\Delta S_{\text{univ}} > 0$ .

**exothermic process** process that gives off heat to the surroundings.

**first law of thermodynamics** energy can be converted from one form to another, but cannot be created or destroyed.

**free energy (G)** energy available to do useful work.

## TIPS FOR THE TEACHER



- heat** form of energy transfer between two bodies at different temperatures.
- heat capacity** quantity of heat required to change the temperature of an object or a given quantity of a substance by one degree Celsius.
- heat engine** device for converting heat into work. The engine absorbs heat at a high temperature, converts part of it to work, and discharges the remaining heat to the surroundings at a lower temperature.
- heat of dilution** heat change (absorption or evolution) associated with the dilution process.
- heat of hydration** heat change (absorption or evolution) associated with the hydration process.
- heat of reaction** quantity of heat associated with a chemical reaction.
- heat of solution** (See enthalpy of solution)
- Hess's law** enthalpy change for an overall or net process is the sum of enthalpy changes for individual steps in the process.
- internal energy** total energy attributed to the particles of matter and their interactions within a system.
- joule (J)** basic SI unit of energy, is the quantity of work done when a force of one newton acts through a distance of one meter.
- kinetic energy** energy available due to the motion of an object.
- molar heat of fusion** energy, expressed in kJ, required to melt one mole of a solid.
- molar heat of sublimation** energy (in kJ) required to sublime one mole of a solid.
- molar heat of vaporization** energy (in kJ) required to vaporize one mole of liquid.
- phase change** transformation from one phase to another.
- potential energy** energy available due to the position of an object.
- second law of thermodynamics** entropy of the universe increases in a spontaneous process and remains unchanged in an equilibrium process.
- specific heat** quantity of heat required to change the temperature of one gram of substance by one degree Celsius.
- state of a system** the values of all pertinent macroscopic variables of a system, *e.g.*, volume, pressure, temperature and composition.
- surroundings** that portion of the universe with which a system interacts.
- system** that portion of the universe selected for a thermodynamic study.
- temperature** measure of the average molecular kinetic energy of a substance.
- thermal energy** energy associated with the random motion of atoms and molecules.
- thermochemical equation** equation that shows both mass and enthalpy relations.
- thermochemistry** the study of the interconversion of heat and other forms of energy in chemical systems.
- third law of thermodynamics** entropy of a pure perfect crystal is zero at the absolute zero for temperature, 0 K.
- work** form of energy transfer that can be expressed as a force acting through a distance.

## Common Student Misconceptions

### 1. “Temperature and heat are the same.”

Students commonly confuse the terms temperature and heat. Temperature is a measure of molecular motion. Heat is a form of energy that flows across a conducting boundary from the higher to the lower temperature. Using units, the difference in these terms is apparent:

- Temperature—°C or K
- Heat—J or kJ (or cal or kcal)

Another example is to have them consider two containers of water, a teacup and a bathtub. Both containers are at the same temperature. Ask students which will cool down fastest. (Obviously, the teacup will be cooled down most quickly, due to the greater amount of heat contained in the bathtub of water relative to that of the teacup.)

Yet another example is to note the similarity and the difference between a burning match and a bonfire. The two temperatures are relatively the same, but the heat content of each is vastly different.

### 2. “Heat capacity, specific heat capacity and molar heat capacity are all the same term.”

These three terms have similar, but different meanings. **Heat capacity** is the amount of heat needed to raise the temperature of a given amount of substance one degree Celsius (or one Kelvin). If 1 gram of the substance is used, the appropriate term for the heat involved is the **specific heat capacity**. For one mole of the substance, the heat involved is the **molar heat capacity**. Unit analysis will help to show the differences:

- Heat capacity—J/C° or J/K
- Specific heat capacity—J/g-C° or J/g-K
- Molar heat capacity—kJ/mol-K or J/mol-K

Two examples that will help students to understand these differences are the following:

- It takes 50.2 J to raise the temperature of 1.0 g of water 12.0 K (1 g x 4.184 J/g-K x 12.0 K); it takes 2510 J to raise the temperature of 50.0 g of water 12.0 K. It takes 904 J to raise the temperature of 1 mole (18 g) of water 12.0 K. Contrast the temperature change, T, and the heat input, q, for each change.
- It takes only 4.78 J to raise the temperature of 1.0 g of copper by 12.0 K; it takes 239 J to raise the temperature of 50.0 g copper by 12.0 K (50.0 g x 0.39 J/g-K x 12.0 K); and it takes 297 J to raise the temperature of 1 mole of copper (63.5 g) 12.0 K. Contrast the temperature change, T, and the heat input, q, for each change.
- Compare and contrast “a” and “b” above.

### 3. “A mechanical engine can be made 100% efficient.” (Automakers could make 100% efficient engines if they wanted to.)

The % efficiency is given by:  $\%E = \frac{T_{\text{hot engine}} - T_{\text{exhaust}}}{T_{\text{hot engine}}} \times 100$

and, in addition, there are always losses due to friction and other factors. Thus, a 100% efficient engine is impossible since there is always a temperature difference between the engine and the exhaust.



4. **“Heat energy can be converted completely into work.”**

Since work can be converted completely into heat, as shown by James Joule and Julius Mayer around 1843, it seems reasonable to assume that heat energy can be converted completely into work, since energy is conserved. This, of course, is **not** the case (No. 3 is an example).

5. **“If a reaction mixture gets hot, the reaction must be endothermic.”**

Apparently the reasoning is as follows: The mixture is hot; therefore, it must have absorbed heat. That makes it an endothermic reaction. Actually, exactly the opposite is true. The reaction is generating heat, and the reaction is exothermic, *i.e.*, heat flows out of the system to heat up the surroundings. This is a very pervasive misconception. Be careful to distinguish between system (loses energy) and surroundings (gains energy and becomes hot).

6. **“The entropy of an element is 0 at 298.2 K and 1 atmosphere.”**

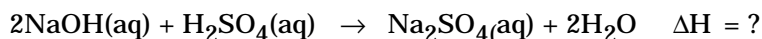
The enthalpy and free energy of the elements at 298.2 K and 1 atmosphere (1 bar) are both 0, but the entropy, an absolute quantity, is not 0 at these conditions. (According to the third law of thermodynamics, the entropy of pure crystalline substances is 0 at 0 K.)

7. **“A calorie must be a huge quantity of energy—look how long it takes to burn off 1 calorie when I exercise.”**

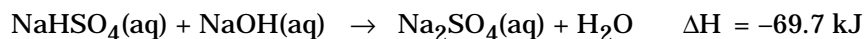
The calorie one burns when exercising is a nutritionist’s calorie, actually 1 kilocalorie or 1000 of a chemist’s calories. The symbol for the food calorie is C (large “C”), representing 1000 calories (small “c”). Remember the calorie was originally defined as the quantity of heat needed to raise the temperature of 1 gram of water 1 C°.

### Problem Solving

1. Calculate the heat of reaction,  $\Delta H$ , for the neutralization of aqueous sulfuric acid by aqueous sodium hydroxide:

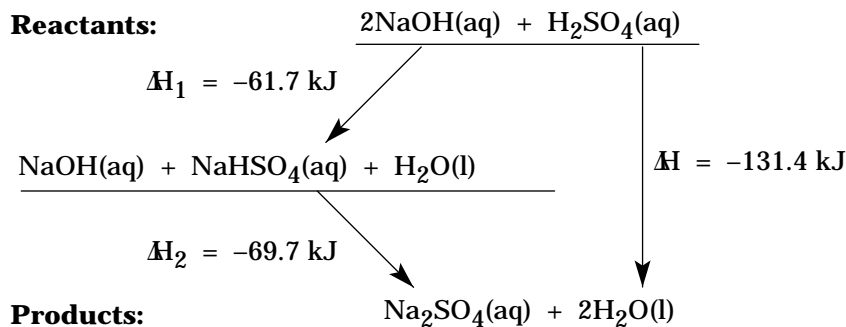


From:

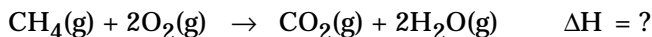


$[\Delta H = -131.4 \text{ kJ}]$

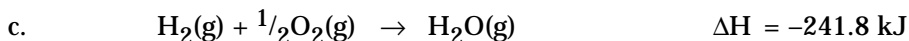
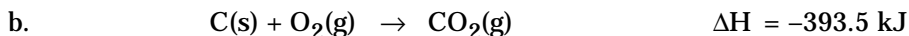
2. Draw an enthalpy diagram to graphically illustrate the heats of the reaction in Question 1.



3. The fuel used for laboratory burners may contain mostly methane,  $\text{CH}_4(\text{g})$ . Calculate the molar heat of combustion of methane, using Hess's Law, from the following data. The equation for the complete combustion of methane is:

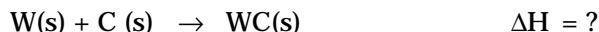


Contributing heat of formation reactions:



[First, reverse Reaction a; next, multiply Reaction c by 2 (for two moles of water formed); then add a, b, and c to get the overall equation, adding  $\Delta H$ 's to obtain  $-802.3 \text{ kJ}$  for the overall reaction.]

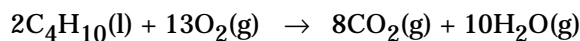
4. Tungsten carbide, WC, is an extremely hard substance commonly found on the cutting surfaces of home and industrial saws and other cutting blades. Calculate the standard heat of formation of tungsten carbide, WC, from its elements, tungsten and graphite (carbon), using the data given below. The equation for the formation of tungsten carbide from its elements is:



	$\Delta H, \text{ kJ}$
$\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$	-393.5
$2\text{WC}(\text{s}) + 5\text{O}_2(\text{g}) \rightarrow 2\text{WO}_3(\text{s}) + 2\text{CO}_2(\text{g})$	-2391.6
$\text{W}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{WO}_3(\text{s})$	-1674.9

[-35.1 kJ/mol WC]

5. Many cigarette lighters contain liquid butane that, when burned, forms  $\text{CO}_2(\text{g})$  and  $\text{H}_2\text{O}(\text{g})$  according to the equation:



Given the following molar heats of formation at standard conditions, calculate:

- the heat of combustion of two moles of butane.
- the heat of combustion of one mole of butane.
- the heat of combustion of 1.0 g butane.
- the heat of reaction for two moles of butane if the product were  $\text{CO}(\text{g})$  instead of  $\text{CO}_2(\text{g})$ .

	$\Delta H_f^\circ$
$\text{C}_4\text{H}_{10}(\text{l})$	-127.4 kJ
$\text{CO}(\text{g})$	-110.5 kJ
$\text{CO}_2(\text{g})$	-393.5 kJ
$\text{H}_2\text{O}(\text{g})$	-241.8 kJ
$\text{H}_2\text{O}(\text{l})$	-285.8 kJ

[a.  $-5311.2 \text{ kJ}$ ; b.  $-2655.6 \text{ kJ}$ ; c.  $-45.7 \text{ kJ}$ ; d.  $-3047.2 \text{ kJ}$ ]



6. When 0.150 g of liquid octane,  $C_8H_{18}(l)$  (molar mass = 114), was burned in a calorimeter containing 1500 g (weighed to the nearest gram) of water, the temperature rose from 24.146 °C to 25.286 °C. Calculate the heat of combustion (kJ/mol) of octane at 25 °C. [The specific heat capacity of water is  $4.184 J g^{-1}C^{-1}$ .]

$$Q = m \times C \times \Delta T$$

$$1500 \text{ g} \times 4.184 \text{ J g}^{-1}C^{-1} \times 1.140 \text{ }^\circ\text{C}$$

$$Q = 7155 \text{ J}$$

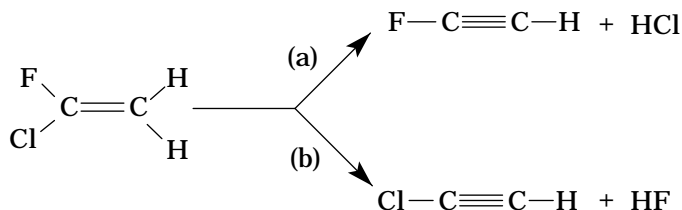
and

$$0.150 \text{ g} \times \frac{1 \text{ mol } C_8H_{18}}{114 \text{ g } C_8H_{18}} = 0.00132 \text{ mol } C_8H_{18}$$

$$\Delta H = -Q_p / \text{mol } C_8H_{18} = \frac{-7155 \text{ J}}{0.00132 \text{ mol } C_8H_{18}} \times \frac{1 \text{ kJ}}{1000 \text{ J}}$$

$$\Delta H = -5420 \text{ kJ/mol } C_8H_{18}$$

7. Kinetic energy may be translational (T), rotational (R), or vibrational (V). Indicate what type(s) of kinetic energy is (are) involved in each of the following:
- Nolan Ryan throws a curve ball. [T, R]
  - Nadja Salerno-Sonnenberg plucks a violin string. [V for string; T for finger]
  - James ("007") Bond plays the bongo drums during a sky diving descent. [T, V, R]
  - The earth moves around the sun. [T, R]
  - A stainless steel ball traverses the deck of a pinball machine. [T, R]
  - A ball attached to a spring is placed in a spinning orbital motion as it hangs from the end of an overhead steel arm. [T, V, R]
  - Steffi Graf executes a backhand at Wimbledon. [T, V, R]
8. Indicate whether the following processes are exothermic (EX) or endothermic (EN) or no heat effect (N).
- A tree grows in the forest. [EN]
  - A tree decays in the forest. [EX]
  - A tree burns in the forest. [EX]
  - Water is formed from the liquid oxygen-liquid hydrogen reaction in the engine of the Discovery rocket. [EX]
  - The water droplets in the rocket exhaust plume evaporate. [EN]
  - The water droplets in the rocket plume freeze. [EX]
  - The water droplets in the rocket plume freeze and then remelt. [N]
  - A cupful of sugar falls to the floor. [EX]
9. A laser may cause either of the following two chemical changes:



Calculate the heat of each reaction,  $\Delta H$ , with the aid of the following average bond dissociation energies (kJ):

Bond	Bond Dissociation Energy (kJ/mol)
C-F	485
C-Cl	339
C-H	413
H-F	565
C=C	607
C≡C	835
C-C	345
H-Cl	431

[Recall that + sign denotes energy added to break a bond while - sign denotes energy given off as a bond forms:

$$\begin{array}{llll}
 \text{a. } & + C=C & + 607 \text{ kJ} & \\
 & + C-Cl & + 339 \text{ kJ} & \\
 & + C-H & + 413 \text{ kJ} & \\
 & & + 1359 \text{ kJ} & \\
 & -C\equiv C & -835 \text{ kJ} & \\
 & -H-Cl & -431 \text{ kJ} & \\
 & & - 1266 \text{ kJ} & \text{net } \Delta H = + 93 \text{ kJ}
 \end{array}$$

$$\begin{array}{llll}
 \text{b. } & + C=C & + 607 \text{ kJ} & \\
 & + C-F & + 485 \text{ kJ} & \\
 & + C-H & + 413 \text{ kJ} & \\
 & & + 1505 \text{ kJ} & \\
 & -C\equiv C & -835 \text{ kJ} & \\
 & -H-F & -565 \text{ kJ} & \\
 & & -1400 \text{ kJ} & \text{net } \Delta H = + 105 \text{ kJ}
 \end{array}$$

10. A typical hamburger has an energy value of 15.0 kJ/g. If a student ate two half-pound hamburgers for lunch (1 lb = 454 g), and if none of the energy was stored in the student's body, how much water must the student sweat off in perspiration so that his body temperature remains constant? The heat of vaporization of water is 2260 J/g. [3.01 kg water]
11. Entropy is related to the number of ways of arranging a system in space, a measure of the degree of disorder. Predict whether the entropy will increase (I) or decrease (D) for each of the following changes:
- Air is pumped into a tire. [D]
  - Gasoline vaporizes in the carburetor of a car. [I]
  - Sugar is dissolved in a cup of coffee. [I]
  - Moisture condenses on the windows of a car. [D]
  - A new deck of playing cards is shuffled. [I]
  - A cylinder of helium is used to fill 3000 balloons for the Thanksgiving Day Parade in New York City. [I]
  - Snowflakes form in the upper troposphere. [D]
  - Magnesium combines with oxygen gas to form magnesium oxide. [D]
  - A box of black jelly beans is mixed with a box of green jelly beans. [I]



## Decision Making

1. Each person, family, community, nation, and the entire world, must develop a program that maximizes energy conservation and minimizes environmental damage. Our present use of fossil fuels is both massive and wasteful. Our future petrochemical industries must have this precious commodity. Scientific, economic, political, and social considerations must be brought to bear on the problem. Available options are suggested below. You may wish to add others. Examine each option in terms of energy needs based upon statistical data and the best up-to-date information and recommend options that strike a proper balance between energy consumption and the reduction of environmental damage. You may wish to restrict your study to your specific locality.
  - a. Accelerate research on nuclear fusion and seek new nuclear power technology.
  - b. Explore increasing the production of synthetic fuels (synfuels) from coal. (Coal gasification and liquefaction are some possible examples.)
  - c. Evaluate the further development of oil from oil shale and tar sands.
  - d. Based on new technology and increased safety factors resulting from European advances in nuclear technology, consider new nuclear fission power plants.
  - e. Improve methodology and engine design that decreases fuel consumption in motor vehicles and aircraft.
  - f. Reconsider the use of animals to replace machines for performing work. (Horses pulling wagons is an example.)
  - g. Set up a public relations program for reducing the amount of energy for lighting, heating, and cooling of area residential and commercial buildings.
  - h. Evaluate one or more of the following alternative energy sources: (a) wind; (b) solar; (c) geothermal; (d) tidal; (e) underground heat pump; (f) ocean thermal.
  - i. Research new energy technology reported in the literature over the past several years.
2. **Energy Consumption in Your Home.** Examine your fuel bills for the past year (or make a yearly projection based on your current monthly bill) to estimate your average monthly and yearly consumption of:
  - a. Electricity in kilowatt hours (1 kWh = 3595 kJ)
  - b. Natural gas in therms or BTU's (1 therm = 1054 kJ)
  - c. Oil in gallonsOn the basis of these estimates for your home, calculate the total average yearly energy consumption in:
  - a. Your town or city
  - b. Your state
  - c. The entire USA
3. **Energy Consumption by Automobiles.** Your car probably runs on gasoline. Determine the weekly consumption of gasoline in gallons by the automobile(s) owned by your family. Now calculate the annual consumption in gallons. Next determine the average yearly consumption of gasoline in:
  - a. Your town
  - b. Your state
  - c. The entire USA

4. A New Source of Crude Oil. A petrologist discovered an underground fossil fuel supply composed of sand, water, and crude oil. Design a series of steps to separate the oil from the mixture. Discuss your proposal with your chemistry teacher. Test your scheme with an artificially prepared mixture in the laboratory. Could you scale up your process so that it could handle large oil spills such as those in the Persian Gulf or off the coast of Alaska?

**Robert Wilhelm Bunsen** (1811-1899), Professor at Kassel, Marburg, and Heidelberg Universities, was known as a great teacher. As a teaching legend at Heidelberg, he taught Adolph Kolbe, Edward Frankland, Victor and Lothar Meyer, Friedrich Beilstein, and Johann Baeyer. Each of these students became a famous chemist in his own right. Bunsen invented the Bunsen burner in 1855 (a similar burner was developed somewhat earlier by Michael Faraday). During research on a highly toxic, unstable compound, he lost an eye during an explosion and twice nearly killed himself through arsenic poisoning. Bunsen and Gustav Kirchhoff invented the spectroscope in 1860 and used it to discover the elements rubidium and cesium.

**Lord Kelvin** (William Thomson), 1824-1907, matriculated at Glasgow University at the age of 10, and then went on to Cambridge. He assimilated and developed for his book on thermodynamics the pioneering principles of James Joule and Nicolas Sadi Carnot. In his 1852 paper "Dissipation of Mechanical Energy," he set down the law of conservation of energy. Kelvin introduced the concept of absolute zero, recognized the theoretical importance of the absolute temperature scale, and calculated the value of absolute zero from theoretical considerations. With Joule, he demonstrated the Joule-Kelvin effect. A famous quotation of Kelvin's is:

"I often say that when you can measure what you are speaking about and express it in numbers, you know something about it, but when you cannot express it in numbers, your knowledge of it is of a meager and unsatisfactory kind; it may be the beginning of knowledge, but you have scarcely, in your thought, advanced to the stage of Science, whatever the latter may be."

In 1840, **James Prescott Joule** (1818-1889) discovered a simple law connecting electrical and heat energy. Stated mathematically,  $E = I^2Rt$ . In 1844, he presented his paper "On the Mechanical Equivalent of Heat" to the Royal Society. Yet both Joule and his father were simple brewers with little formal education. The son carried out his laboratory work in a private laboratory built by his father near their brewery.

Germain **Henri Hess** (1802-1850), although born in Switzerland, was taken to Russia as a child. First a physician, he became a professor of chemistry at the Technological Institute of the University of St. Petersburg. In 1840, by carefully measuring the heats of reaction for many chemical changes, he was able to develop his Law of Constant Heat Summation, known today simply as Hess's Law. The law states that in any chemical reaction, regardless of how many steps there are, the quantity of heat developed in the overall reaction is constant.

**Josiah Willard Gibbs** (1839-1903) was considered a poor teacher but a brilliant and productive theorist. His early interest in the theory behind James Watt's steam engine governor was responsible for his later interest in thermodynamics. In a series of papers published between 1873 and 1876, he developed virtually all the basic theoretical principles of thermodynamics. His paper "Equilibrium of Heterogeneous Substances" contains the widely used Gibbs "phase rule." In addition, Gibbs developed the vitally important concept of free energy, symbolized by  $\Delta G$ , that predicts reaction direction.

## HISTORY: ON THE HUMAN SIDE



## HUMOR: ON THE FUN SIDE

1. From CHEM 13 NEWS, January 1977, p. 144:

A calorie-counter named Newell  
Was confounded for want of a rule,  
Till someone implied,  
If by 4.184 multiplied,  
Each calorie'd turn to a joule.

2. Word Search (see *Appendix* for master copy)

```
H Z B C I T E N I K Y S R T P J H
G E Y U W Y E H K P H R C C V K N
A I J K R B I X O R M L J T W X K
V R E T E M I R O L A C P K G Y N
I O N J $ L T X K T W X S U D V H
H L C K $ N N E Y U H B E U U N K
Z A M X E E N V A X Y E C K O D Q
M C V E H H E S O J X K R U W V N
R C I C Q J H X Y H Q O K M K E G
E T J L N M M D Z P W U C G I N O
X D Y Z F C H F R Z Z H L Z Y C Z
U P S X E R U T A R E P M E T L Y
U T H E R M O C H E M I S T R Y Y
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. System's measure of disorder or randomness.
2. Device to measure the heat quantity exchanged between a system and its surroundings.
3. SI unit of energy.
4. Measure of the average molecular kinetic energy of a substance.
5. Formulator of a fundamental thermochemical law.
6. Study of the interconversion of heat and other forms of energy in chemical systems.
7. Describing a process that gives off heat to the surroundings.
8. Energy of a system by virtue of its motion.
9. Force acting through a distance.
10. Quantity of heat needed to raise the temperature of one gram of water 1 degree Celsius.

Answers: 1. ENTROPY 2. CALORIMETER 3. JOULE 4. TEMPERATURE  
5. HESS 6. THERMOCHEMISTRY 7. EXOTHERMIC 8. KINETIC 9. WORK  
10. CALORIE

3. See cartoons at end of the module.

## MEDIA

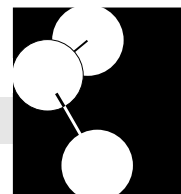
1. *The World of Chemistry* videotape "Number 8: Chemical Bonds" discusses bond strength and energy involvement; "Number 13: The Driving Forces" discusses why reactions occur using heats of reaction. *World of Chemistry Videocassettes*. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; *World of*

- Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
2. CHEM Study films/videos available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9012; (800) 962-2660.
    - a. *Molecular Motions*. Through animation of molecules, this film illustrates vibrational, rotational and translational kinetic energies. It covers phase changes and simple chemical changes resulting in changes in these three types of energy.
    - b. *Vibration of Molecules*. This film describes different vibrational modes in simple molecules. Some of the content may be too difficult for the average college preparatory class.
    - c. *Gas Pressure and Molecular Collisions*. This film describes effects of increasing temperature and pressure on gaseous systems using animated models. It is useful in helping students visualize kinetic energy (temperature) and motion of molecules.
  3. "The Laws of Disorder," Imperial Chemical Industries, Ltd., ICI Film Library, Thames House North, Millbank, London SW1P 4QG.
  4. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396; (608) 262-5153 (voice) or (608) 262-0381 (FAX).
    - a. *MolVib: Visualizing Molecular Vibrations*, by Daniel Huber. Vol. III C, No. 1, for the Apple Macintosh.
    - b. *Microstate*, by Richard York. Vol. V B, No. 2, for the IBM PS/2 PC-compatible computers.
    - c. *The Computer-Based Laboratory*, by Daniel Krause. Vol. I A, No. 2, for the Apple II computer.
    - d. *Bomb Calorimeter Simulation*, by David J. Olney. Vol. III A, No. 2, for the Apple II computer.
    - e. *Notebook: A Column Calculator and Plotter*, by Robert Rittenhouse. Vol. IV B, No. 1, for IBM PS/2 PC-compatible computers.
  5. Videodiscs published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396; (608) 262-5153 (voice) or (608) 262-0381 (FAX).
    - a. "An Exothermic Reaction" and "An Endothermic Reaction," two chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.). Special Issue 3.
    - b. "Thermite Reaction," a chapter on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.). Special Issue 4.
  6. *Redox*, by Helen Brooks and David Brooks. Videodisc distributed by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396; (608) 263-2837; (608) 262-0381 (FAX).

The following materials from Project SERAPHIM can be used to build a thermistor probe. Laboratory modules LM 003, LM 004, LM 005, LM 006, LM 015, LM 016, LM 021 (for use with Apple II computers) and LM 022 (for use with IBM PC-compatible computers); LM 023. Published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396; (608) 263-2837; (608) 262-0381 (FAX).

## EQUIPMENT AND INSTRUMENTATION

# Links/Connections



## WITHIN CHEMISTRY

1. Molecular motion (translational, rotational, vibrational) and temperature
2. Vibrational kinetic motion and infrared spectroscopy
3. Heats of reaction and potential energy diagrams (used extensively in the *Reaction Rates* module)

## BETWEEN CHEMISTRY AND OTHER DISCIPLINES

### Physics

1. Potential energy, stored energy, heat energy.
2. Gravitational potential energy of a heavy textbook on the floor, work needed to lift it to a table, a position of greater potential energy—all directly analogous to changes in heat energy during a chemical reaction.
3. Heat vs. temperature.
4. Heat flow from warmer to cooler surroundings.
5. Calorimetry experiments, especially specific heat determinations of various substances.

### Earth Science

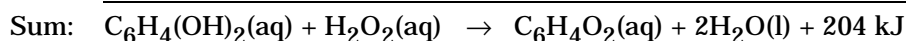
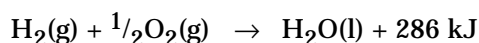
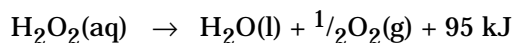
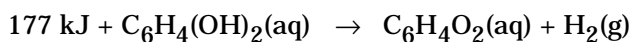
1. Snowballs or rocks rolling downhill, losing potential energy as they are converted to kinetic and heat energies.
2. Freezing and/or evaporating of water and the concomitant heat changes.
3. Geothermal energy sources within the earth.

### Biology

1. Maintenance of body temperature by combustion reactions within the body.
2. The bombardier beetle provides interesting examples of both exothermic reactions and Hess's Law when it attempts to protect itself from its natural enemies. A pair of glands at the tip of its abdomen has an inner compartment containing hydroquinone,  $C_6H_4(OH)_2(aq)$ , and hydrogen peroxide,  $H_2O_2$ , which are squeezed into an outer compartment containing enzymes (catalysts). A rapid, exothermic reaction occurs, producing quinone,  $C_6H_4O_2$ , and water, and releasing 204 kJ of heat, according to the equation:

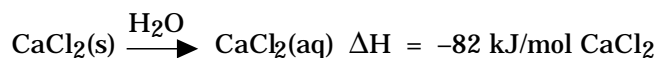


The heat raises the solution to its boiling point and the vapor thus produced is sprayed out up to 30 times in quick succession as a hot fine mist with audible sound effects. The Hess's Law example emerges when the above reaction is broken down into its three separate components:

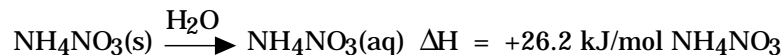


**Personal**

1. Chemical reactions involving the burning of foods maintain our body temperature and make us fat if we are not careful about what and how much we eat.
2. Attention we all pay to food calories (kilocalories) in our daily diet.
3. Combustion of fuels to warm our homes in winter.
4. Gasoline mileage in our cars related to efficiency of automobile engines.
5. Hot packs and cold packs (used to treat athletic injuries) are the result of chemical reactions also.

**Hot pack:**

Just 40 g  $\text{CaCl}_2$  mixed with 100 g water will raise the temperature of the water from 20 °C to 90 °C.

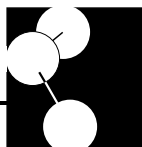
**Cold pack:**

Only 30 g ammonium nitrate mixed with 100 g water will lower the temperature of the water from 20 °C to 0 °C.

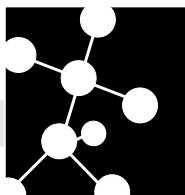
**Community**

1. A field trip to a local power generating plant may be educational, whether a coal-burning, oil-burning, or nuclear plant.
  2. Guest speakers from a power plant may be instructive.
- 
1. Energy shortages, oil embargo, unfriendly takeover of oil-producing countries, oil wars.
  2. Environmental problems caused by our energy needs:
    - a. Oil spills
    - b. Air pollution, acid rain
    - c. Greenhouse effect
  3. The energy needed to propel the space shuttle into orbit is equivalent to the energy produced by 23 Hoover dams. The shuttle engines, consisting of three main engines and the two side solid booster rockets, consume 11,000 lb of fuel each second. The total thrust is approximately 44 million horsepower, which is equivalent to 400,000 subcompact cars or 157,000 six-axle diesel locomotives.
  4. The solar energy received by the Earth annually is equivalent to 46 times the initial world reserves of coal or 425 times the initial world reserves of crude oil. Only 0.04% of this solar energy is involved in photosynthesis, involving 200 billion tons of  $\text{CO}_2$  and releasing 400 billion tons of oxygen.

**TO THE  
CONTEMPORARY  
WORLD****SOCIETAL  
(SCIENCE/  
TECHNOLOGY/  
SOCIETY;  
CURRENT  
EVENTS)**



5. On the last page of his 1923 textbook on thermodynamics, G. N. Lewis expressed concern about the negative  $\Delta G$  value for the decomposition of  $\text{NH}_4\text{NO}_3$ , especially noting that a mole of solid ammonium nitrate upon heating will decompose into three moles of gas. No one seems to have paid much attention at the time. Ammonium nitrate was a widely used fertilizer, and the storage of large bags filled with it, both in barns and in warehouses, was commonplace. About 25 years later in Texas City, where bags of ammonium nitrate fertilizer were being unloaded on the docks, a fire broke out on one of the ships, and the ship exploded. The fire then spread, and the resulting series of explosions blew up much of Texas City. About 600 people were killed, and almost 4,000 others were injured.



# Extensions

## 1. Heats of Reaction involving Acids and Bases

- Measure 100 mL of 0.500 M NaOH into a double polystyrene cup. (Nest one cup inside the other.) Place 100 mL of 0.500 M HCl into a beaker. Note and record the initial temperature ( $^{\circ}\text{C}$ ) of each solution. Rinse and dry both thermometer and graduated cylinder after each use. Take an average of these two initial temperatures as the temperature of the mixture before reaction.
- Mix the two solutions with gentle stirring. With the use of a  $0.1^{\circ}\text{C}$  (or a  $1^{\circ}\text{C}$ ) thermometer, read and record the highest temperature reached as the final temperature, and then calculate the change in temperature,  $\Delta T$ . Next calculate the amount of heat produced, in joules, assuming the solution has a heat capacity of  $4.184 \text{ J/g } ^{\circ}\text{C}$ . Finally, calculate the quantity of heat produced when you mix one mole of each reactant together.
- Repeat the above procedure, using other acid-base combinations: (1)  $\text{HNO}_3 - \text{NaOH}$ ; (2)  $\text{HCl} - \text{KOH}$ ; (3)  $\text{HNO}_3 - \text{KOH}$ ; (4)  $\text{H}_2\text{SO}_4 - \text{NaOH}$ ; (5)  $\text{CH}_3\text{COOH} - \text{NaOH}$ ; (6)  $\text{CH}_3\text{COOH} - \text{NH}_3$ . What conclusions about the nature of each of these reactions can you make, based on your data?

## 2. Heats of Solution of Various Salts

- Prepare two sets of double polystyrene cups, and place 100 mL of water in each. Obtain separately 3.0 g  $\text{CaCl}_2 \cdot 6 \text{H}_2\text{O}$  and 6.0 g  $\text{NH}_4\text{Cl}$  (or  $\text{NH}_4\text{NO}_3$ ). Record the initial temperature of the water in each calorimeter to the nearest  $0.1^{\circ}\text{C}$  (more accurate temperature measurement is desirable). Add the  $\text{CaCl}_2 \cdot 6 \text{H}_2\text{O}$  into the first calorimeter, stir or swirl, and record the highest temperature reached. Repeat with the  $\text{NH}_4\text{Cl}$  in the other calorimeter, recording the lowest temperature reached.
- Use the above data to calculate the heat change and heat change per mole, both for the  $\text{CaCl}_2 \cdot 6 \text{H}_2\text{O}$  and the  $\text{NH}_4\text{Cl}$ . Repeat with other salts. Repeat using methanol and 95% ethanol in place of the water.

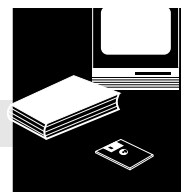
## 3. A Chemist's Map of Pennsylvania

Note the energy-related communities on the Pennsylvania map (see *Appendix*): Energy, Boiling Springs, Steam Valley, Crucible, Gastown, Coalport, Coal Hill, and Oil City. Select another state from a detailed map atlas and make a list and/or map of all the sites with names related to energy. Choose one or more of the energy-related sites and try to discover the historical background of the name(s).

## 4. Thermal Pollution: A Library Research Project

An important environmental concern with regard to nuclear power is thermal pollution. Cooling towers and holding reservoirs for hot water in nuclear plants impact the aqueous environment. Adding hot water to a body of water causes irreversible damage to the organisms of this habitat, killing some flora and fauna and increasing eutrophication. (Idea contributed by Angie Matamoros.)

# References



Module drafted by Herbert Bassow, William Bleam, and Arthur C. Breyer, the Pennsylvania team.

Barnwell, G. (1990, December). Your own personal greenhouse. *ChemMatters*, 8(4), 8-10.

A brief overview of how our energy requirements contribute to greenhouse gases. A good chemistry-community reference.

Bartlett, A. (1978). Forgotten fundamentals of the energy crisis. *American Journal of Physics*, 49, 867-888.

Benson, S. (1965). Bond energies. *Journal of Chemical Education*, 42, 502-518.

Importance of bond energy concept in field of chemistry. Includes many tables of bond energies and strengths.

Bent, H. A. (1967). *The second law*. New York, NY: Oxford University Press.

Benzinger, T. H. (1961, January). The human thermostat. *Scientific American*, 204(1), 134-147.

Dyson, F. J. (1954, September). What is heat? *Scientific American*, 191(3), 58-63.

Knox, B. E., and Palmer, H. B. (1961). The use and misuse of bond energies. *Journal of Chemical Education*, 38, 292-294.

MacWood, G. E., and Verhoek, F. H. (1961). How can you tell whether a reaction will occur? *Journal of Chemical Education*, 38, 334-337.

Marsella, G. (1987). Hot and cold packs. *ChemMatters*, 5(1), 6-11.

An overview of chemistry and applications of thermal packs with a follow up do-it-yourself experiment.

McGlashen, M. L. (1966). The use and misuse of the laws of thermodynamics. *Journal of Chemical Education*, 43, 226-232.

What the field of thermodynamics is and what it isn't.

Parry, R. W., Bassow, H., and Merrill, P. (1987). *Chemistry: Experimental foundations* (Teacher's Guide, pp. 381-382). Englewood Cliffs, NJ: Prentice-Hall.

This is the source for the suggested demonstration on entropy increase during an endothermic reaction.

Porter, G. (1968a). The laws of disorder. *Chemistry*, 41(5), 23-25.

Porter, G. (1968b). The laws of disorder. *Chemistry*, 41(6), 21-24.

Porter, G. (1968c). The laws of disorder. *Chemistry*, 41(7), 36-37.

Porter, G. (1968d). The laws of disorder. *Chemistry*, 41(8), 16.

Porter, G. (1968e). The laws of disorder. *Chemistry*, 41(9), 28.

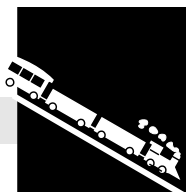
Porter, G. (1968f). The laws of disorder. *Chemistry*, 41(10), 24.

Porter, G. (1968g). The laws of disorder. *Chemistry*, 41(11), 28.

- Porter, G. (1968h). The laws of disorder. *Chemistry*, 42(1), 21.
- Porter, G. (1968i). The laws of disorder. *Chemistry*, 42(2), 19.
- Russo, T. (1990). *Micro chemistry lab manual* (Vol. 2, 3). WestChester, OH: Kemtec Educational Corp., 9889 Crescent Park Drive.
- Micro chemistry experiments for high school chemistry. Volumes 2 and 3 contain experiments on thermochemistry and calorimetry.
- Sanderson, R. T. (1968). Why does methane burn? *Journal of Chemical Education*, 45, 423-425.
- Summerlin, L. R., and Ealy, J. (1985). *Chemical demonstrations* (Vol. 1). Washington, DC: American Chemical Society.
- Swinton, F. L. (1967). The triple point of water. *Journal of Chemical Education*, 44, 541.
- Tripp, T. B. (1976). The definition of heat. *Journal of Chemical Education*, 53, 782-784.
- Weast, R. (1971) *Handbook of chemistry and physics* (52nd Ed.). Cleveland, OH: Chemical Rubber Company.
- The table of covalent bond strengths is taken from pages F177-184.
- The Laws of Disorder is a series of brief articles that parallel a 10-part series on chemical change and thermodynamics presented by Nobel Laureate Sir George Porter, *Chemistry of Very Fast Reactions* (1967), for the BBC.
- |              |  |
|--------------|--|
| 41(5), 23-25 | Introduction to the Series               |
| 41(6), 21-24 | The First Law of Thermodynamics          |
| 41(7), 36-37 | Spontaneity and the Second Law           |
| 41(8), 16    | Equilibrium: The Limit of Disorder       |
| 41(9), 28    | Thermal Efficiency and Useful Work       |
| 41(10), 24   | Factors Affecting Reaction Rates         |
| 41(11), 28   | Reaction Order and Collision Probability |
| 42(1), 21    | How Light Initiates Reactions            |
| 42(2), 19    | Series Conclusion: Chaos and Evolution   |
- A good essay for students trying to resolve science and religion.

NOTE: This series is said to have included 100 demonstrations.

# Appendix



- **Transparency Masters**
  1. Table of Single Covalent Bond Energies (from *Handbook of Chemistry and Physics*)
  2. Enthalpy Changes during Reactions
  3. Schematic Diagram of a Calorimeter
  4. Heats of Reactions between Elements
  5. Enthalpy Changes in the Decomposition and Formation of Water
  6. Dissolving of Solid Ammonium Nitrate ( $\text{NH}_4\text{NO}_3$ ): A Typical Endothermic Process
  7. A Chemist's Map of Pennsylvania
  8. Word Search
  
- **Humor**

### Single Covalent Bond Energies

$\text{kJ mol}^{-1}$  ( $\text{kcal mol}^{-1}$ )

	<b>F</b>	<b>O</b>	<b>N</b>	<b>Cl</b>	<b>Br</b>	<b>I</b>	<b>C</b>	<b>H</b>	<b>Si</b>
<b>Si</b>	582 (139)	368 (87.9)	335 (80.1)	391 (93)	310 (74)	234 (55)	301 (71.9)	323 (77)	226 (54)
<b>H</b>	566 (135)	467 (111.6)	391 (93)	431 (103)	366 (87.4)	299 (71.5)	416 (99)	436 (104.2)	
<b>C</b>	485 (115.9)	336 (80.3)	285 (68)	327 (78)	285 (68)	213 (50.9)	356 (85)		
<b>I</b>	192.5 (46)	201 (48)		209 (49.9)	180 (43)	151 (36.1)			
<b>Br</b>	238 (56.8)			217 (51.9)	193 (46.1)				
<b>Cl</b>	255 (60.9)	205 (49)	193 (46)	241.7 (57.8)					
<b>N</b>	272 (65)	201 (48)	160 (38)						
<b>O</b>	190 (45)	146 (34.9)							
<b>F</b>	158 (37.8)								

#### Additional Bond Energies

$\text{O}=\text{O}$  498.7  $\text{kJ mol}^{-1}$  (118.66  $\text{kcal mol}^{-1}$ )

$\text{N}\equiv\text{N}$  941.4  $\text{kJ mol}^{-1}$  (226.80  $\text{kcal mol}^{-1}$ )

$\text{P}_2$  489  $\text{kJ mol}^{-1}$  (116.70  $\text{kcal mol}^{-1}$ )

$\text{C}=\text{C}$  620  $\text{kJ mol}^{-1}$

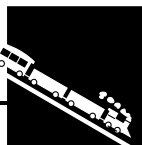
$\text{C}\equiv\text{C}$  812  $\text{kJ mol}^{-1}$

$\text{C}=\text{N}$  615  $\text{kJ mol}^{-1}$

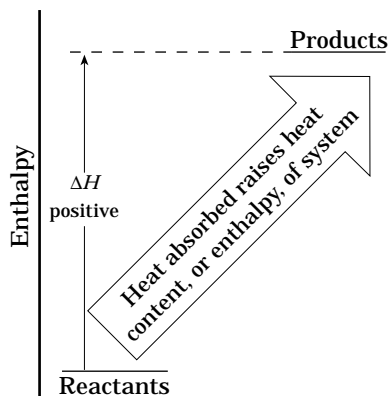
$\text{C}\equiv\text{N}$  891  $\text{kJ mol}^{-1}$

$\text{C}=\text{O}$  781  $\text{kJ mol}^{-1}$

$\text{N}=\text{N}$  418  $\text{kJ mol}^{-1}$

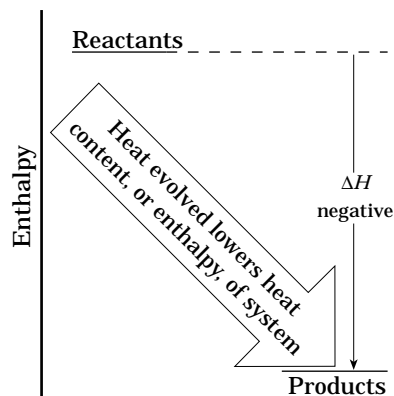


## Enthalpy Changes during Reactions



Reaction proceeding →  
Energy is absorbed  
 $\Delta H$  is positive

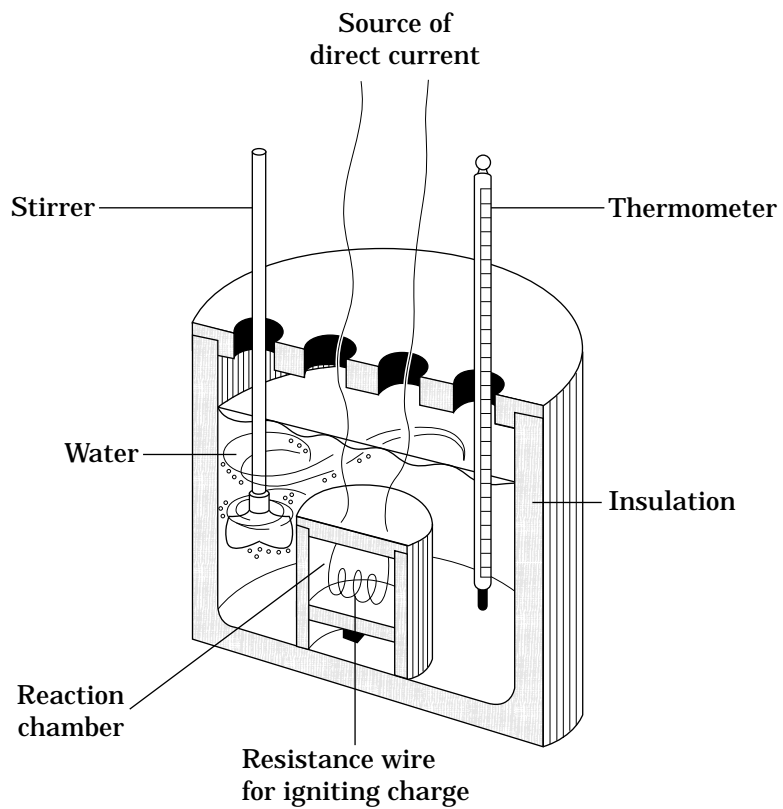
(a)

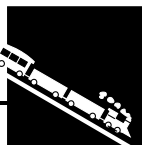


Reaction proceeding →  
Energy is evolved  
 $\Delta H$  is negative

(b)

### Schematic Diagram of a Calorimeter

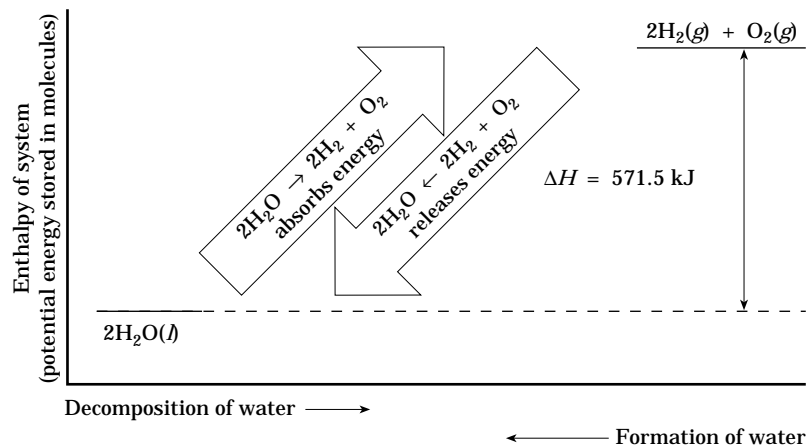


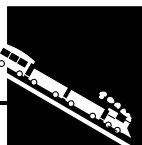


## Heats of Reactions between Elements

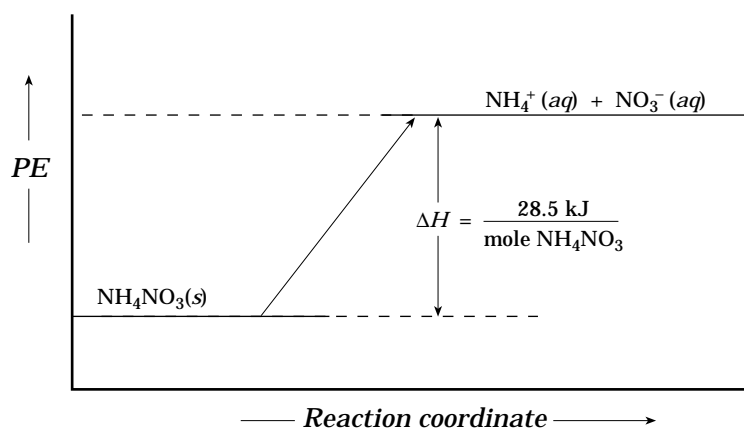
Elements	Compound (product)	Heat of reaction* (kJ/mole of product) $\Delta H$
$\text{H}_2(g) + \frac{1}{2} \text{O}_2(g)$	$\longrightarrow \text{H}_2\text{O}(g)$   Water vapor	-241.8
$\text{H}_2(g) + \frac{1}{2} \text{O}_2(g)$	$\longrightarrow \text{H}_2\text{O}(l)$   Water	-285.8
$\text{S}(s) + \text{O}_2(g)$	$\longrightarrow \text{SO}_2(g)$   Sulfur dioxide	-297.0
$\text{H}_2(g) + \text{S}(s) + 2\text{O}_2(g)$	$\longrightarrow \text{H}_2\text{SO}_4(l)$   Sulfuric acid	-811.7
$\frac{1}{2} \text{N}_2(g) + \frac{1}{2} \text{O}_2(g)$	$\longrightarrow \text{NO}(g)$   Nitric oxide	90.4
$\frac{1}{2} \text{N}_2(g) + \text{O}_2(g)$	$\longrightarrow \text{NO}_2(g)$   Nitrogen dioxide	33.9
$\frac{1}{2} \text{N}_2(g) + \frac{3}{2} \text{H}_2(g)$	$\longrightarrow \text{NH}_3(g)$   Ammonia	-46.3
$\text{C}(s) + \frac{1}{2} \text{O}_2(g)$	$\longrightarrow \text{CO}(g)$   Carbon monoxide	-110.5
$\text{C}(s) + \text{O}_2(g)$	$\longrightarrow \text{CO}_2(g)$   Carbon dioxide	-393.3
$\text{C}(s) + 2\text{H}_2(g)$	$\longrightarrow \text{CH}_4(g)$   Methane	-74.5
$2\text{C}(s) + 3\text{H}_2(g)$	$\longrightarrow \text{C}_2\text{H}_6(g)$   Ethane	-84.5
$3\text{C}(s) + 4\text{H}_2(g)$	$\longrightarrow \text{C}_3\text{H}_8(g)$   Propane	-103.8
$\frac{1}{2} \text{H}_2(g) + \frac{1}{2} \text{I}_2(g)$	$\longrightarrow \text{HI}(g)$   Hydrogen iodide	25.9
$\text{S}(s) + \frac{3}{2} \text{O}_2(g)$	$\longrightarrow \text{SO}_3(g)$   Sulfur trioxide	-394.6

\* T = 25 °C, P = 1 atm

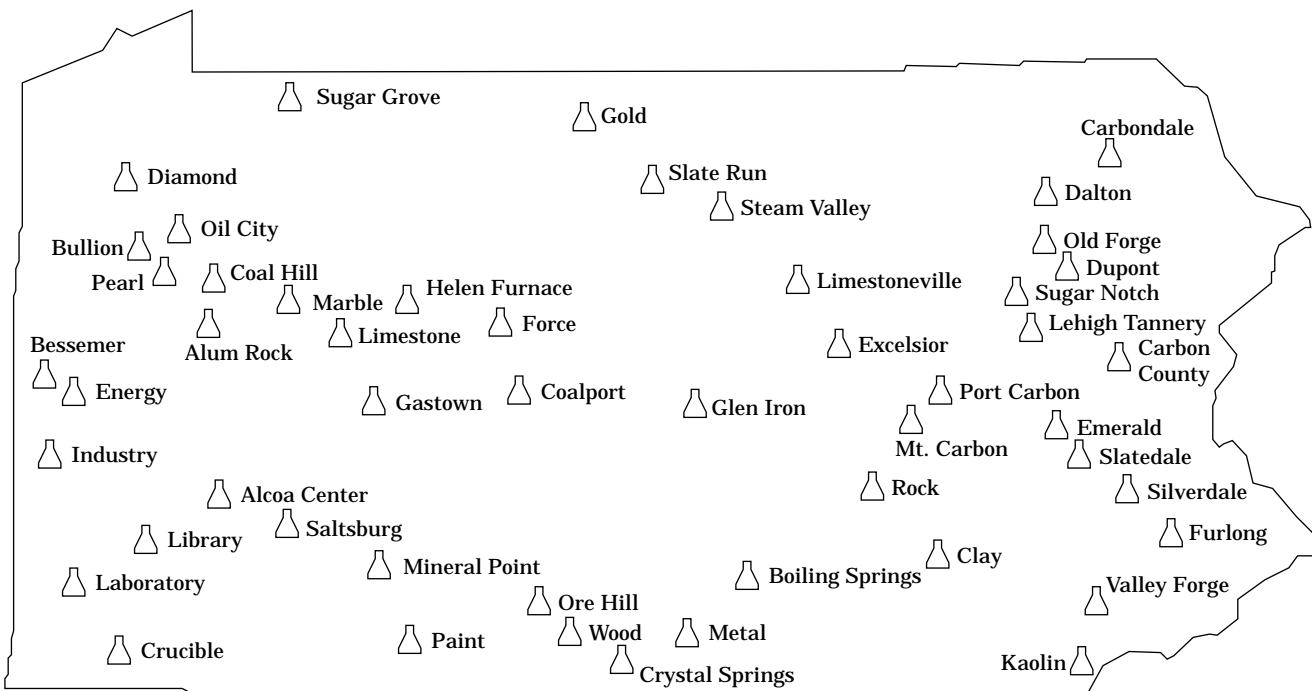
**Enthalpy Changes in the Decomposition and Formation of Water**



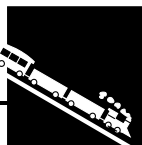
## Dissolving of Solid Ammonium Nitrate ( $\text{NH}_4\text{NO}_3$ ): A Typical Endothermic Process



## A Chemist's Map of Pennsylvania



Courtesy of J. J. Bohning, Professor of Chemistry Emeritus, Wilkes University, and Assistant Director, Chemical Heritage Foundation, 3401 Walnut Street, Philadelphia, PA 19104-6228.



## Word Search

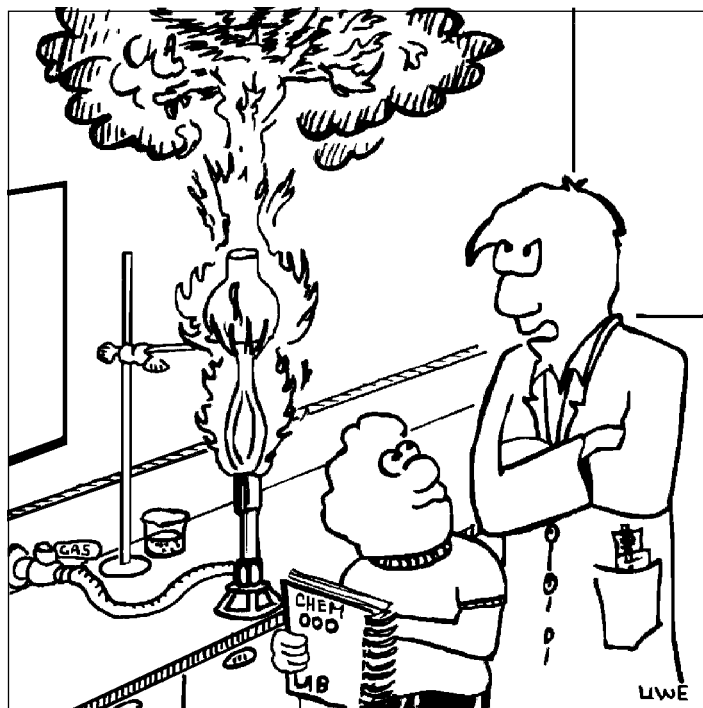
H Z B C I T E N I K Y S R T P J H  
G E Y U W Y E H K P H R C C V K N  
A I J K R B I X O R M L J T W X K  
V R E T E M I R O L A C P K G Y N  
I O N J S L T X K T W X S U D V H  
H L C K S N N E Y U H B E U U N K  
Z A M X E E N V A X Y E C K O D Q  
M C V E H H E S O J X K R U W V N  
R C I C Q J H X Y H O O K M K E G  
E T J L N M M D Z P W U C G I N O  
X D Y Z F C H F R Z Z H L Z Y C Z  
U P S X E R U T A R E P M E T L Y  
U T H E R M O C H E M I S T R Y Y

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. System's measure of disorder or randomness.
2. Device to measure the heat quantity exchanged between a system and its surroundings.
3. SI unit of energy.
4. Measure of the average molecular kinetic energy of a substance.
5. Formulator of a fundamental thermochemical law.
6. Study of the interconversion of heat and other forms of energy in chemical systems.
7. Describing a process that gives off heat to the surroundings.
8. Energy of a system by virtue of its motion.
9. Force acting through a distance.
10. Quantity of heat needed to raise the temperature of one gram of water 1 degree Celsius.

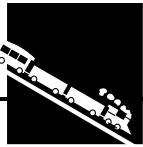


Joyce M Edson, Softwear Unlimited International, P.O. Box 265, Twisp, WA 98856. T-shirts and posters are available from this company. Reprinted with permission.



**YES WILLY, I KNOW IT SAYS TO HEAT BUT!**

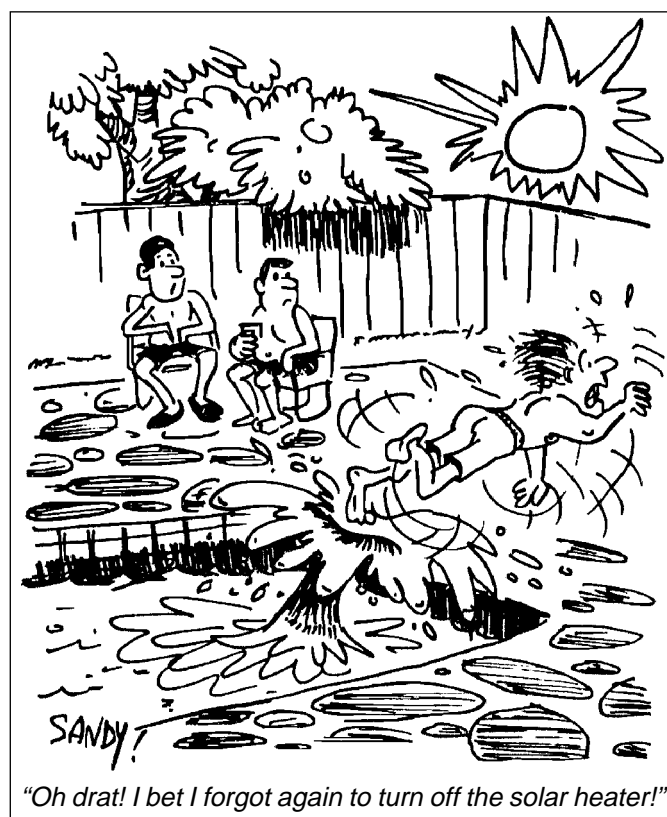
CHEM 13 NEWS, November 1983, p. 3. Reprinted with permission.



© American Chemical Society. CHEMTECH, November 1973, p. 704. Reprinted with permission.



Cartoon by Shelly Fischman. Reprinted with permission.



© American Chemical Society. CHEMTECH, August 1984, p. 472. Reprinted with permission.