

Chapter 5. Thermochemistry

Media Resources

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- Figure 5.10 Internal Energy Is a State Function, but Heat and Work Are Not
- Figure 5.13 Pressure–Volume Work
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- 5.3 Enthalpy

Section:

- 5.2 The First Law of Thermodynamics
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- 5.8 Fuels and Foods

Other Resources

Further Readings:

Weight-Loss Diets and the Law of Conservation of Energy
 Pictorial Analogies III: Heat Flow, Thermodynamics, and Entropy
 Analogical Demonstrations
 Heat Flow vs. Cash Flow: A Banking Analogy
 Three Forms of Energy
 A Specific Heat Analogy
 Heat Capacity, Body Temperature, and Hypothermia
 Calories—Who's Counting?
 The Conversion of Chemical Energy: Part 1. Technological Examples
 The Geochemistry of Coal. Part II: The Components of Coal
Scientific American, September 1990
 Hydrogen: The Ultimate Fuel and Energy Carrier
 Chemical Fuels from the Sun
 "The Ice that Burns. Can Methane Hydrates Fuel the 21st Century?

Section:

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 5.3 Enthalpy
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 5.8 Foods and Fuels
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Live Demonstrations:

Evaporation as an Endothermic Process
 Flaming Cotton
 Heat of Neutralization
 Chemical Cold Pack
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 Endothermic Reaction: Ammonium Nitrate
 Boiling Water in a Paper Cup: Heat Capacity of Water
 Making Canned Heat

Section:

5.2 The First Law of Thermodynamics
 5.4 Enthalpies of Reaction
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Chapter 5. Thermochemistry

Common Student Misconceptions

- Students confuse power and energy.
- Students confuse heat with temperature.
- Students fail to note that the first law of thermodynamics *is* the law of conservation of energy.
- Students have difficulty in determining what constitutes the system and the surroundings.
- Sign conventions in thermodynamics are always problematic.
- Students do not realize that a chemical reaction carried out in an open container occurs at constant pressure.
- Since enthalpy is a state function, and $\Delta H = q_p$, students often think that *heat* is a state function.
- Students often cannot tell the difference between enthalpy of a reaction ΔH_{rxn} (in kJ) and molar enthalpy (per one mole of one of the reacting species, in kJ/mol).
- Students do not realize that Hess's law is a consequence of the fact that enthalpy is a state function.
- Students should be directed to Appendix C of the text for a list of standard enthalpy values. (They are unlikely to find this information on their own!)
- Students tend to have difficulties with calculating a value of ΔH°_f for a compound not listed in Appendix C from ΔH°_{rxn} and the available ΔH°_f values.
- Students tend to think that any reaction in which a given compound is formed, regardless of the type of reactants, should be called a *formation reaction*.
- Students often neglect to notice that a *formation reaction* leads to a formation of *1 mole* of a compound.

Teaching Tips

- Remind students that the values of ΔH°_f for the same compound but in a different phase are different; this is one of main causes of errors in calculations in section 5.7.
- Emphasize that molar heat capacity and specific heat capacity have different units; one is expressed on a per-mole basis while the other is expressed on a per-gram basis.

Lecture Outline

5.1 The Nature of Energy

- **Thermodynamics** is the study of energy and its transformations.
- **Thermochemistry** is the study of the relationships between chemical reactions and energy changes involving heat.
- Definitions:
 - **Energy** is the capacity to do work or to transfer heat.
 - **Work** is energy used to cause an object with mass to move.

$$w = F \times d$$
 - **Heat** is the energy used to cause the temperature of an object to increase.
 - A **force** is any kind of push or pull exerted on an object.
 - The most familiar force is the pull of gravity.

Kinetic Energy and Potential Energy

- **Kinetic energy** is the energy of motion:

$$E_k = \frac{1}{2}mv^2$$

- **Potential energy** is the energy an object possesses by virtue of its position or composition.
 - Electrostatic energy is an example.
 - It arises from interactions between charged particles.

$$E_{el} = \frac{\kappa Q_1 Q_2}{d}$$

- Potential energy can be converted into kinetic energy.
 - An example is a ball of clay dropped off a building.

Units of Energy

- SI unit is the **joule** (J).
- From $E_k = \frac{1}{2}mv^2$, $1\text{J} = 1\text{kg} \times \frac{\text{m}^2}{\text{s}^2}$
- Traditionally, we use the **calorie** as a unit of energy.
 - $1\text{ cal} = 4.184\text{ J}$ (exactly)
- The nutritional Calorie, Cal = $1,000\text{ cal} = 1\text{ kcal}$.

System and Surroundings

- A **system** is the part of the universe we are interested in studying.
- **Surroundings** are the rest of the universe (i.e., the surroundings are the portions of the universe that are not involved in the system).
- Example: If we are interested in the interaction between hydrogen and oxygen in a cylinder, then the H₂ and O₂ in the cylinder form a system.

Transferring Energy: Work and Heat

- From physics:
 - **Force** is a push or pull exerted on an object.
 - **Work** is the energy used to move an object against a force.
 $w = F \times d$
 - **Heat** is the energy transferred from a hotter object to a colder one.
 - **Energy** is the capacity to do work or to transfer heat.

FORWARD REFERENCES

- The concepts of the system, surroundings, and universe will be used again in Chapter 19.
- Energy in J will be used in Chapter 6 to express energy of a hydrogen atom and the energy of an emitted photon.
- Energy in kJ/mol will be used in Chapter 7 for ionization energies and electron affinities.
- Energy in kJ/mol will be used in Chapter 8 to measure lattice energy and average bond enthalpies and to calculate enthalpies of reactions.
- Energy in J (and kJ) to energy of thermodynamic calculations throughout Chapters 19, 14 (Arrhenius equation in section 14.5), and Chapter 20 (Gibbs free energy vs. cell potential in section 20.5).

5.2 The First Law of Thermodynamics¹

- **The first law of thermodynamics** states that energy cannot be created or destroyed.
- The first law of thermodynamics is the law of conservation of energy.
 - That is, the energy of system + surroundings is constant.
 - Thus, any energy transferred from a system must be transferred to the surroundings (and vice versa).

¹ “Weight-Loss Diets and the Law of Conservation of Energy” from Further Readings

Internal Energy²

- The total energy, E , of a system is called the **internal energy**.
 - It is the sum of all the kinetic and potential energies of all components of the system.
 - Absolute internal energy cannot be measured, only changes in internal energy.
 - Change in internal energy: $\Delta E = E_{\text{final}} - E_{\text{initial}}$.
 - Thermodynamic quantities such as ΔE always have three parts:
 - a number,
 - a unit, and
 - a sign that gives direction.
 - Example: A mixture of $\text{H}_2(g)$ and $\text{O}_2(g)$ has a higher internal energy than $\text{H}_2\text{O}(g)$.
 - Going from $\text{H}_2(g)$ and $\text{O}_2(g)$ to $\text{H}_2\text{O}(g)$ results in a negative change in internal energy, indicating that the system has lost energy to the surroundings:
- $$\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g) \quad \Delta E < 0$$
- Going from $\text{H}_2\text{O}(g)$ to $\text{H}_2(g)$ and $\text{O}_2(g)$ results in a positive change in internal energy, indicating that the system has gained energy from the surroundings:
- $$2\text{H}_2\text{O} \rightarrow \text{H}_2(g) + \text{O}_2(g) \quad \Delta E > 0$$

Relating ΔE to Heat and Work³

- From the first law of thermodynamics:
 - When a system undergoes a physical or chemical change, the change in internal energy is given by the heat added to or liberated from the system plus the work done on or by the system:
$$\Delta E = q + w$$
- Heat flowing from the surroundings to the system is positive, $q > 0$.
- Work done by the surroundings on the system is positive, $w > 0$.

Endothermic and Exothermic Processes

- An **endothermic** process is one that *absorbs* heat from the surroundings.
 - An endothermic reaction feels cold.
- An **exothermic** process is one that *transfers* heat to the surroundings.
 - An exothermic reaction feels hot.

State Functions^{4,5,6}

- A **state function** depends only on the initial and final states of a system.
 - Example: The altitude difference between Denver and Chicago does not depend on whether you fly or drive, only on the elevation of the two cities above sea level.
 - Similarly, the internal energy of 50 g of $\text{H}_2\text{O}(l)$ at 25 °C does not depend on whether we cool 50 g of $\text{H}_2\text{O}(l)$ from 100 °C to 25 °C or heat 50 g of $\text{H}_2\text{O}(l)$ at 0 °C to 25 °C.
- A state function does not depend on how the internal energy is used.
 - Example: A battery in a flashlight can be discharged by producing heat and light. The same battery in a toy car is used to produce heat and work. The change in internal energy of the battery is the same in both cases.

FORWARD REFERENCES

- Equilibria of endothermic vs. exothermic reactions will be differently affected by temperature changes in Chapter 15.
- Water autoionization is mentioned to be an endothermic process (Chapter 16, section 16.3).

² Figure 5.5 from Transparency Pack

³ “Sign Conventions for q and w ” Activity from Instructor’s Resource CD/DVD

⁴ “Thermite” Movie from Instructor’s Resource CD/DVD

⁵ “Evaporation as an Endothermic Process” from Live Demonstrations

⁶ Figure 5.10 from Transparency Pack

- Solid → liquid → gas phase changes are endothermic, while gas → liquid → solid phase changes are exothermic.
- Most spontaneous reactions (Chapter 19) are exothermic.
- State functions will be further discussed in Chapter 19 (entropy and Gibbs free energy).

5.3 Enthalpy^{7,8,9,10,11}

- Chemical and physical changes that occur around us occur under essentially constant pressure of Earth's atmosphere.
 - Changes may be accompanied by work done by or on the system.
 - Changes may involve the release or absorption of heat.
 - We will focus much of our discussion on what we can learn from measurement of heat flow.
- **Enthalpy (H)** is defined as the internal energy (E) plus the product of the pressure and volume of the system.

$$H = E + PV$$

- Enthalpy is useful for investigating heat flow in events that occur under constant pressure.
- Again, we can only measure the change in enthalpy, ΔH .
- Mathematically,

$$\begin{aligned}\Delta H &= H_{\text{final}} - H_{\text{initial}} = \Delta E + P\Delta V \\ w &= -P\Delta V; \Delta E = q + w \\ \Delta H &= \Delta E + P\Delta V = (q_p + w) - w = q_p\end{aligned}$$

- For most reactions $P\Delta V$ is small thus $\Delta H = \Delta E$
- Heat transferred from surroundings to the system has a positive enthalpy (i.e., $\Delta H > 0$ for an endothermic reaction).
- Heat transferred from the system to the surroundings has a negative enthalpy (i.e., $\Delta H < 0$ for an exothermic reaction).
- Enthalpy is a state function.

A Closer Look at Energy, Enthalpy, and P-V Work¹²

- Many chemical reactions involve work done on or by the system.
 - Work is often either electrical or mechanical work.
 - Mechanical work done by a system involving expanding gases is called **pressure-volume work** or **P-V work**.
- Consider:
 - A cylinder of cross-sectional area A ,
 - A piston exerting a pressure, $P = F/A$, on a gas inside the cylinder,
 - The volume of gas expanding through ΔV while the piston moves a height $\Delta h = h_f - h_i$.
 - The magnitude of work done = $F \times \Delta h = P \times A \times \Delta h = P \times \Delta V$.
 - Since work is being done by the system on the surroundings,
 - $w = -P\Delta V$.
 - Using the first law of thermodynamics,
 - $\Delta E = q - P\Delta V$.
 - If the reaction is carried out under constant volume,
 - $\Delta V = 0$ and $\Delta E = q_v$.

⁷ "Heat Flow vs. Cash Flow: A Banking Analogy" from Further Readings

⁸ "Pictorial Analogies III: Heat Flow, Thermodynamics, and Entropy" from Further Readings

⁹ "Analogical Demonstrations" from Further Readings

¹⁰ "Work of Gas Expansion" Animation from Instructor's Resource CD/DVD

¹¹ "Three Forms of Energy" from Further Readings

¹² Figure 5.13 from Transparency Pack

- If the reaction is carried out under constant pressure,
 - $\Delta E = q_p - P\Delta V$, or
 - $q_p = \Delta H = \Delta E + P\Delta V$
 - and $\Delta E = \Delta H - P\Delta V$

5.4 Enthalpies of Reaction^{13,14,15,16,17,18,19,20}

- For a reaction, $\Delta H_{rxn} = H_{products} - H_{reactants}$.
- The enthalpy change that accompanies a reaction is called the **enthalpy of reaction** or *heat of reaction* (ΔH_{rxn}).
- Consider the thermochemical equation for the production of water:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g) \quad \Delta H = -483.6 \text{ kJ}$$
 - The equation tells us that 483.6 kJ of energy are released to the surroundings when water is formed.
 - ΔH noted at the end of the balanced equation depends on the number of moles of reactants and products associated with the ΔH value.
 - These equations are called *thermochemical equations*.
- *Enthalpy diagrams* are used to represent enthalpy changes associated with a reaction.
- In the enthalpy diagram for the combustion of $H_2(g)$, the reactants, $2H_2(g) + O_2(g)$, have a higher enthalpy than the products $2H_2O(g)$; this reaction is exothermic.
- *Enthalpy is an extensive property*.
 - Therefore, the *magnitude* of enthalpy is directly proportional to the amount of reactant consumed.
 - Example: If one mol of CH_4 is burned in oxygen to produce CO_2 and water, 890 kJ of heat is released to the surroundings. If two mol of CH_4 is burned, then 1780 kJ of heat is released.
- The *sign* of ΔH depends on the direction of the reaction.
 - The enthalpy change for a reaction is equal in magnitude but opposite in sign to ΔH for the reverse reaction.
 - Example: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -890 \text{ kJ}$,
 - But $CO_2(g) + 2H_2O(l) \rightarrow CH_4(g) + 2O_2(g) \quad \Delta H = +890 \text{ kJ}$.
- Enthalpy change depends on *state of the products and reactants*.
 - $2H_2O(g) \rightarrow 2H_2O(l) \quad \Delta H = -88 \text{ kJ}$

FORWARD REFERENCES

- Enthalpies of chemical reactions and physical processes will be used throughout the textbook.
- Enthalpies of reactions will be calculated using average bond enthalpies in Chapter 8.
- Enthalpy changes will be used in Chapter 19 to evaluate Gibbs free energy changes.

5.5 Calorimetry

- **Calorimetry** is a measurement of heat flow.
- A **calorimeter** is an apparatus that measures heat flow.

¹³ “Enthalpy of Solution” Activity from Instructor’s Resource CD/DVD

¹⁴ “Formation of Water” Movie from Instructor’s Resource CD/DVD

¹⁵ “Flaming Cotton” from Live Demonstrations

¹⁶ “Heat of Neutralization” from Live Demonstrations

¹⁷ “Chemical Cold Pack” from Live Demonstrations

¹⁸ “A Chemical Hand Warmer” from Live Demonstrations

¹⁹ “Endothermic Reaction: Ammonium Nitrate” from Live Demonstrations

²⁰ “Beware-Fertilizer Can EXPLODE!” from Further Readings

Heat Capacity and Specific Heat²¹

- **Heat capacity** is the amount of energy required to raise the temperature of an object by 1 °C.
 - **Molar heat capacity** is the heat capacity of 1 mol of a substance.
 - **Specific heat**, or specific heat capacity, is the heat capacity of 1 g of a substance.
- Heat, $q = (\text{specific heat}) \times (\text{grams of substance}) \times \Delta T$.
- Be careful with the sign of q .

Constant-Pressure Calorimetry^{22,23}

- The most common technique is to use atmospheric pressure as the constant pressure.
- Recall $\Delta H = q_p$.
- The easiest method is to use a coffee cup calorimeter.

$$q_{\text{soln}} = (\text{specific heat of solution}) \times (\text{grams of solution}) \times \Delta T = -q_{\text{rxn}}$$
- For dilute aqueous solutions, the specific heat of the solution will be close to that of pure water.

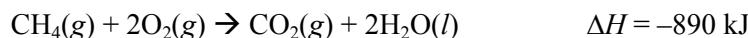
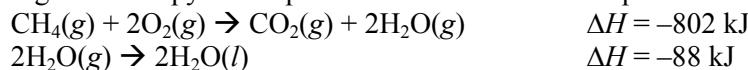
Bomb Calorimetry (Constant-Volume Calorimetry)²⁴

- Reactions can be carried out under conditions of constant volume instead of constant pressure.
- Constant volume calorimetry is carried out in a **bomb calorimeter**.
- The most common type of reaction studied under these conditions is combustion.
- If we know the heat capacity of the calorimeter, C_{cal} , then the heat of reaction,

$$q_{\text{rxn}} = -C_{\text{cal}} \times \Delta T$$
- Since the reaction is carried out under constant volume, q relates to ΔE .

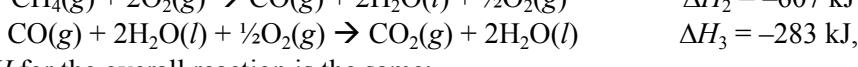
5.6 Hess's Law^{25,26}

- **Hess's Law:** If a reaction is carried out in a series of steps, ΔH for the reaction is the sum of ΔH for each of the steps.
- The total change in enthalpy is independent of the number of steps.



- Therefore, for the reaction $\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(l)$, $\Delta H = -890 \text{ kJ}$.
- Note that ΔH is sensitive to the states of the reactants and products.
- Total ΔH is also independent of the nature of the path.

- If we convert $\text{CH}_4(g) + 2\text{O}_2(g)$ to $\text{CO}_2(g) + 2\text{H}_2\text{O}(l)$ through a CO intermediate:



Then ΔH for the overall reaction is the same:

$$\Delta H_2 + \Delta H_3 = -607 \text{ kJ} - 283 \text{ kJ} = -890 \text{ kJ} = \Delta H_1$$

FORWARD REFERENCES

- Similar rules will apply to ΔG and ΔS in Chapter 19.

²¹ “Boiling Water in a Paper Cup: Heat Capacity of Water” from Live Demonstrations

²² “A Specific Heat Analogy” from Further Readings

²³ “Calorimetry” Activity from Instructor’s Resource CD/DVD

²⁴ Figure 5.19 from Transparency Pack

²⁵ “Heat Capacity, Body Temperature, and Hypothermia” from Further Readings

²⁶ Figure 5.22 from Transparency Pack

5.7 Enthalpies of Formation^{27,28}

- If a compound is formed from its constituent elements, then the enthalpy change for the reaction is called the **enthalpy of formation**, ΔH_f .
- Standard state* (standard conditions) refer to the substance at:
 - 1 atm and 25 °C (298 K).
- Standard enthalpy**, ΔH° , is the enthalpy measured when everything is in its standard state.
- Standard enthalpy of formation** of a compound, ΔH_f° , is the enthalpy change for *the formation of 1 mol of compound* with all substances in their *standard states*.
- ΔH_f° for selected substances are tabulated in Appendix C.
 - A large majority of ΔH_f° values tabulated in Appendix C are negative, meaning that most formation reactions are exothermic.
- If there is more than one state for a substance under standard conditions, the more stable one is used.
Example: When dealing with carbon we use graphite because graphite is more stable than diamond or C_{60} .
- The standard enthalpy of formation of the most stable form of an element is zero.

Using Enthalpies of Formation to Calculate Enthalpies of Reaction^{29,30}

- Use Hess's law!
 - Example: Calculate ΔH for

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$
 - We start with the reactants, decompose them into elements, then rearrange the elements to form products. The overall enthalpy change is the sum of the enthalpy changes for each step.
 - Decomposing into elements (note O_2 is already elemental, so we concern ourselves with C_3H_8):
 $C_3H_8(g) \rightarrow 3C(s) + 4H_2(g)$ $\Delta H_1 = -\Delta H_f^\circ [C_3H_8(g)]$
 - Next we form CO_2 and H_2O from their elements:
 $3C(s) + 3O_2(g) \rightarrow 3CO_2(g)$ $\Delta H_2 = 3 \Delta H_f^\circ [CO_2(g)]$
 $4H_2(g) + 2O_2(g) \rightarrow 4H_2O(l)$ $\Delta H_3 = 4 \Delta H_f^\circ [H_2O(l)]$
 - We look up the values and add:
 $\Delta H_{rxn}^\circ = -1(-103.85 \text{ kJ}) + 3(-393.5 \text{ kJ}) + 4(-285.8 \text{ kJ}) = -2220 \text{ kJ}$
 - In general:
- $$\Delta H_{rxn}^\circ = \Delta H_f^\circ(\text{products}) - \Delta H_f^\circ(\text{reactants})$$
- Where n and m are the stoichiometric coefficients.

FORWARD REFERENCES

- In Chapter 19 (section 19.5) Gibbs free energies of formation, ΔG_f° , will be used to find ΔG_{rxn}° in an analogous manner to how ΔH_f° values are used to find ΔH_{rxn}° .

²⁷ “Formation of Aluminum Bromide” Movie from Instructor’s Resource CD/DVD

²⁸ Table 5.3 from Transparency Pack

²⁹ “Nitrogen Triiodide” Movie from Instructor’s Resource CD/DVD

³⁰ Figure 5.23 from Transparency Pack

5.8 Foods and Fuels

- **Fuel value** is the energy released when 1 g of substance is burned.
- The fuel value of any food or fuel is a positive value that must be measured by calorimetry.

Foods^{31,32,33,34,35,36}

- Fuel value is usually measured in Calories (1 nutritional Calorie, 1 Cal = 1000 cal).
 - Most energy in our bodies comes from the oxidation of carbohydrates and fats.
 - In the intestines carbohydrates are converted into glucose, C₆H₁₂O₆, or blood sugar.
 - In the cells glucose reacts with O₂ in a series of steps, which produce CO₂, H₂O, and energy.
- $$\text{C}_6\text{H}_{12}\text{O}_6(s) + 6\text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 6\text{H}_2\text{O}(l) \quad \Delta H^\circ = -2803 \text{ kJ}$$
- Fats, for example tristearin, react with O₂ as follows:
- $$2\text{C}_{57}\text{H}_{110}\text{O}_6(s) + 163\text{O}_2(g) \rightarrow 114\text{CO}_2(g) + 110\text{H}_2\text{O}(l) \quad \Delta H^\circ = -75,250 \text{ kJ}$$
- Fats contain more energy than carbohydrates. Fats are not water soluble. Therefore, fats are good for energy storage.

Fuels^{37,38,39,40,41,42,43,44}

- In 2008 the United States consumed about 1.05×10^{17} kJ/year (9.4×10^5 kJ per person per day).
- Most of this energy comes from petroleum and natural gas.
- The remainder of the energy comes from coal, nuclear and hydroelectric sources.
- Coal, petroleum, and natural gas are **fossil fuels**. They are not renewable.
- **Natural gas** consists largely of carbon and hydrogen. Compounds such as CH₄, C₂H₆, C₃H₈ and C₄H₁₀ are typical constituents.
- **Petroleum** is a liquid consisting of hundreds of compounds. Impurities include S, N, and O compounds.
- **Coal** contains high molecular weight compounds of C and H. In addition compounds containing S, O, and N are present as impurities that form air pollutants when burned in air.

Other Energy Sources⁴⁵

- **Nuclear energy**: energy released in splitting or fusion of nuclei of atoms.
 - It is used to produce about 21% of the electric power in the US.
- Fossil fuels and nuclear energy are *nonrenewable* sources of energy.
- **Renewable energy** sources include:

³¹ “Calories—Who’s Counting?” from Further Readings

³² “Sucrose” 3-D Model from Instructor’s Resource CD/DVD

³³ “Carbon Dioxide” 3-D Model from Instructor’s Resource CD/DVD

³⁴ “Oxygen” 3-D Model from Instructor’s Resource CD/DVD

³⁵ “Water” 3-D Model from Instructor’s Resource CD/DVD

³⁶ “Ethanol” 3-D Model from Instructor’s Resource CD/DVD

³⁷ “Making Canned Heat” from Live Demonstrations

³⁸ Figure 5.25 from Transparency Pack

³⁹ “The Conversion of Chemical Energy. Part 1. Technological Examples” from Further Readings

⁴⁰ “The Geochemistry of Coal. Part II. The Components of Coal” from Further Readings

⁴¹ *Scientific American, September, 1990* from Further Readings

⁴² “Hydrogen: The Ultimate Fuel and Energy Carrier” from Further Readings

⁴³ “Chemical Fuels from the Sun” from Further Readings

⁴⁴ “The Ice that Burns. Can Methane Hydrates Fuel the 21st Century?” from Further Readings

⁴⁵ “Hydrogen” 3-D Model from Instructor’s Resource CD/DVD

- Solar energy
- Wind energy
- Geothermal energy
- Hydroelectric energy
- Biomass energy
- These are virtually inexhaustible and will become increasingly important as fossil fuels are depleted.

FORWARD REFERENCES

- Pollution resulting from fossil-fuel combustion will be discussed in Chapter 18 (section 18.2).
- Pollution-free fuel cell-powered vehicles will be mentioned in Chapter 20 (section 20.7).

Further Readings:

1. John J. Fortman, "Pictorial Analogies III: Heat Flow, Thermodynamics, and Entropy," *J. Chem. Educ.*, Vol. 70, **1993**, 102–103.
2. John J. Fortman, "Analogical Demonstrations," *J. Chem. Educ.*, Vol. 69, **1992**, 323–324. This reference contains a quick analogical demonstration on heat transfer.
3. John W. Hill, "Weight-Loss Diets and the Law of Conservation of Energy," *J. Chem. Educ.*, Vol. 58, **1981**, 996.
4. Charles M. Wynn, Sr., "Heat Flow vs. Cash Flow: A Banking Analogy," *J. Chem. Educ.*, Vol. 74, **1997**, 3978.
5. Sigthor Petursson, "Three Forms of Energy," *J. Chem. Educ.*, Vol. 80, **2003**, 776–778.
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