Chemistry and the Life Science Standards

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Students have a genuine interest in biological processes, as they are relevant to their daily lives. Students also have first-hand experience with many biological concepts. Therefore, using biological examples in chemistry classrooms can not only tap into students' inherent interest in biology, but also help them make important interdisciplinary connections. Additionally, *No Child Left Behind* and state science standards expect *all* students to have a more comprehensive and connected understanding of science.

However, the *National Science Education Standards* (NSES) suggest that many students associate molecules with physical science and fail to understand that living systems are composed of molecules. The inability of students to make connections between important biological processes and key chemistry concepts such as atoms and molecules, structure and properties of matter, thermodynamics, and chemical reactions is largely the result of how we teach chemistry. We want our students to understand that chemistry is relevant to their everyday lives, yet many instructional examples involve chemicals or contexts with which students have no practical experience. This makes it difficult for students to construct the desired connections to other disciplines and encourages them to confine chemistry concepts to the chemistry classroom. Moreover, research indicates the transfer of information from one context to another is enhanced by presentation of material in multiple contexts (Bransford, 2000).

Therefore, it is notable that in the NSES, there is not listed any chemistry standards. Although most teachers clearly recognize that many of the physical science standards relate specifically to chemistry, there are also a number of life science standards that incorporate chemistry concepts. The overlap between the chemistry concepts found in the physical and life science standards is shown in Table 1.

Table 1: Chemistry in the Life Science and Physical Science Standards

Physical Science Standards	Related Life Science Standards
Structure of atoms Structure and properties of matter Chemical reactions Conservation of energy	**The Cell** Cells are made up of molecules Different types of molecules form specialized cell structures (e.g., membranes) Most cell functions involve chemical reactions Cells use chemical reactions to produce energy
Structure and properties of matter	**Molecular basis of heredity** Chemical and structural properties of DNA explain how genetic information is encoded and replicated
Conservation of energy and increase in disorder Chemical reactions Interactions of matter and energy	**Matter, energy, and organization in living systems** All matter tends toward more disorganized states Plants capture energy from the sun to form covalent chemical bonds ATP is a high-energy molecule

Table 2 summarizes changes that will promote the integration of chemistry and the life sciences. The remainder of the chapter discusses each content area in Table 2 in detail, including annotated curricular materials and highlighted classroom examples. This does not require teaching a separate course in organic chemistry or biochemistry, but rather linking certain key chemistry concepts to life science teaching.

Table 2: Changing Emphases To Integrate Chemistry and the Life Sciences

Content Area	Less emphasis on	More emphasis on
Structure and shape	primarily inorganic examples VSEPR to teach shape with no discussion of context	Lewis structures of organic molecules and organic functional groups biological importance of molecular shape
Intermolecular forces and physical properties	small-molecule examples of intermolecular forces	biological examples of intermolecular forces
Chemical reactions	reactions and problems involving reagents unfamiliar to students	reaction and stoichiometry problems with biological significance and small organic molecules (such as acetic acid)
Thermodynamics	calorimetry examples with metals and water Hess's law examples with inorganic reactions	food metabolism examples and labs biological examples of exothermic and endothermic reactions Hess's law examples with small organic molecules

Structure and Shape

Generally, lessons on ionic and covalent bonding are followed by a discussion of the structure and shapes of molecules. Students typically learn about the octet rule and how to draw Lewis structures, as well as how to use Lewis structures to determine the three-dimensional shapes of molecules using Valence Shell Electron Pair Repulsion (VSEPR) theory. Traditionally, we use small inorganic molecules to teach these concepts because students can focus on just a few atoms. However, the majority of covalent compounds that students will encounter in their lives are organic molecules; the structure and shape of organic molecules play an important role in many biological processes.

One example in which the structure and shape play an important role is in smell. The structure of organic molecules impacts smell as different organic functional groups tend to have different types of odors. For example, amines generally have a fishy smell, esters have a sweet fruity smell, and carboxylic acids have a pungent or putrid smell. Figure 1 shows two organic molecules with the same molecular formula but different functional groups. These two compounds have very different odors.

Figure 1. Molecules with identical composition, but different structure.

Ethyl acetate-sweet banana smell
Found in rum flavored extract

Butyric acid-putrid smell of rancid butter
Can obtain from Adrich chemical

The *Living by Chemistry* smells curriculum (Stacy et al., 2003*a*) uses smells to introduce students to organic molecules. By classifying compounds based on smell and comparing the molecular formulae of similar smelling compounds, students are able to relate organic functional groups to particular smells. Furthermore, by examining a number of organic structures, students begin to develop the rules for drawing organic structures: H atoms form one bond, O atoms 2, N atoms 3, and C atoms 4. They call this the “HONC—1234 rule.” Connections can then be made to the periodic table, with the number of valance electrons providing the basis for drawing Lewis structures of both organic and inorganic molecules.

However, it is not just the functional group that determines the smell of a compound; the shape of a molecule also affects its odor. Similar to the way that substrate molecules fit into enzyme-active sites like a key fits into a lock, smell receptors also bind molecules of a specific shape. Discovering the relationship between odor and shape can lead to a discussion of VSEPR and three-dimensional shape. Students can begin by making models of small molecules such as CH_4, NH_3, H_2O, CO_2, C_2H_6, CH_3OH, etc., to clearly see how the number of bonds and lone pairs around the central atom affect the three-dimensional shapes of molecules. Finally, students can either build or examine prebuilt models of larger organic molecules to make connections between shape and smell. For example, sweet-smelling compounds tend to be linear or “stringy,” flat molecules have a minty odor, and more spherical molecules smell like camphor (Stacy et al., 2003*b*). Organic molecules provide excellent examples of molecular shapes up to tetrahedral. Going beyond tetrahedral electron pair geometries is usually unnecessary, as students will not encounter molecules with 5 or 6 regions of electrons around the central atom until more advanced college chemistry courses.

How Sweet It Is

Purpose: To demonstrate the effect of molecular shape on taste

Background: Most students are familiar with the concept of photosynthesis—plants make glucose from CO_2, H_2O, and light—but few have any real understanding of glucose on the molecular level. Additionally, students may not know that there are different kinds of sugars and that sugar units (monosaccharides) link together to form larger sugars or carbohydrates (Institute of Food Technologies, 2007). Glucose and fructose are both sugars with the chemical formula $C_6H_{12}O_6$; however, fructose is much sweeter than glucose. Sucrose is used as the standard for measuring sweetness and is given the value of 1. Fructose has a sweetness value of 170 and glucose 80 (Ontario Ministry of Agriculture, 2007). The sweetness of a sugar is largely dependent on its shape (Emsley, 1988; Purves, 2006).

Activity: After students have learned about VSEPR and three-dimensional shape, have them build models of fructose and glucose to examine how two molecules with the same molecular formula can have different three-dimensional shapes. Students can also connect the models of glucose and fructose to see how sucrose (table sugar) is formed. Structures can be found on the Web (Lyndaker, 2007).

Assessment: Students can be asked the following questions: (1) What is similar and what is different about the structures of glucose and fructose? (2) How could these structural differences affect how sweet it tastes? (3) What is removed when you connect glucose and fructose together to form sucrose?

Extensions: Students can build a model of an artificial sweetener, aspartame, and compare its structure to sucrose and fructose. The structure of aspartame can be found at http://www.3dchem.com.

Intermolecular Forces and Physical Properties

Typically, after molecular shape, students learn about polarity and intermolecular forces, as well as how intermolecular forces impact physical properties, such as melting point and solubility. However, we often forget that intermolecular forces play a very important role in the life sciences. For example, polarity plays an important role in vitamin solubility. Polar vitamins (like vitamin C) are water soluble, whereas nonpolar vitamins (like vitamin E) are fat soluble. An example of the important role that lipids (nonpolar molecules) play in the body is found in the *ChemMatters* student activity "The Aspirin Effect: Pain Relief and More" (Kimbrough, 2004). In reading this article, students learn about the role of prostaglandins in pain and swelling and how drugs like aspirin, ibuprofen, and acetaminophen (Tylenol) work to inhibit the prostaglandin synthesis pathway. An important connection to chemistry is that aspirin and ibuprofen work similarly and contain similar functional groups, but acetaminophen has different functional groups and inhibits a different part of the synthesis pathway. The teacher's guide for this activity provides additional instructional materials, including background information, student assessment questions, and structured student reading guides, as well as a table connecting each activity to the NSES.

Noncovalent interactions are also important in DNA and protein structure. The DNA double helix is held together by hydrogen bonds; adenine always pairs with thymine and cytosine with guanine because these pairs have maximal hydrogen bonding interactions within the confined geometry of the double helix. Hydrogen bonds between amino acid amide groups are also what hold protein alpha helices and beta sheets together. The three-dimensional folding of proteins depends on both hydrogen bonding and nonpolar interactions. Often, the nonpolar amino acid side chains are found on the inside of the protein, forming a hydrophobic core, and the polar groups are found on the outside of the protein, helping to make the protein soluble in water. Polar and nonpolar interactions are also important for enzyme substrate specificity. Both the shape of the active site and noncovalent interactions between the amino acid side chains in the active site and the substrate ensure that only specific molecules will fit into an active site. Molecular structures of DNA, DNA base pairs, alpha helices, beta sheets, and small proteins can be found on the Web (Ophardt, 2003).

More important than just placing a chemistry concept into a biological context, using biological examples may help address a common student misconception. Despite emphasizing the difference between covalent bonds and intermolecular forces and stressing that when water boils or sugar dissolves, bonds are not being broken, many students hold onto the belief that for a liquid to enter the gas phase or when a covalent compound dissolves in water, it must break apart. In fact, some entering chemistry graduate students still have this misconception (Bodner, 1991). However, challenging this misconception with a biological example can help students develop the correct scientific explanation. *Living by Chemistry* again uses the concept of smells to do this (Stacy et al., 2003*b*). Students can understand that although the substances they smell are liquids, to travel to their nose through the air, they must become gases. To address the question of whether molecules break up or remain whole when they leave the liquid state and become a gas, students are presented with four pictorial models: (1) atoms in the liquid state and atoms in the gas state, (2) molecules in the liquid state and molecules in the gas state, (3) atoms in the liquid state and molecules in the gas state, and (4) molecules in the liquid state and atoms in the gas state. Students are then guided through an activity comparing butyric acid and ethyl acetate (see Fig. 1), testing the models to see which best accounts for the observed physical data. This strategy is similar to that of the Process-Oriented Guided Inquiry Learning activities described in chapter 4 of this book.

Chemical Reactions and Stoichiometry

The study of chemical reactions usually focuses first on the qualitative evidence of a chemical reaction (color change, gas evolution, temperature change, precipitate formation, etc.) followed by writing balanced chemical equations for reactions; categorizing reactions into four or five types (synthesis, decomposition, single replacement, double replacement, and combustion of hydrocarbons); and finally, introducing coefficients in chemical equations as mole ratios (stoichiometry). As inorganic compounds contain relatively few atoms or ions, they are ideally suited for this core qualitative and quantitative approach to chemical reactions that is central to chemistry. However, inorganic reactions often use elements and compounds that are unfamiliar to students. Using biologically relevant examples can provide students with a familiar context as well as draw upon their natural interest in biology.

Focusing on particular inorganic substances that affect (or may be found in) the human body and how quantities of these substances are measured provides a more biologically focused approach to studying chemical reactions. Many inorganic ions and compounds have biological connections, such as calcium phosphate (found in bones), lead ions (toxic to humans), and calcium oxalate (one type of kidney stone). Common blood tests such as the CHEM-20 include tests for serum (blood) calcium and chloride ions. These ions are precipitated with oxalate and silver ions, respectively. These examples can be used in stoichiometry problems and laboratory experiments. An online medical encyclopedia contains normal levels of these ions in healthy human blood to provide a realistic and health-focused context for the problems (Medline Plus, 2005). Addressing stoichiometry problems involving serum can also be used to introduce and practice net ionic equations, as well as solution stoichiometry. The ions and compounds mentioned above can be seamlessly added to precipitation labs and problem sets. Here is an example of how a straightforward stoichiometry exercise can be transformed into a problem with a biological context:

Old exercise: If 1.50 g of magnesium iodide reacts with an excess of sodium carbonate, what mass of magnesium carbonate will be formed, according to the equation below?

$$MgI_2\,(aq) + Na_2CO_3\,(aq) \rightarrow MgCO_3\,(s) + 2NaI\,(aq)$$

Kool-Aid Chromatography

Purpose: To use chromatography to separate food dyes based on polarity

Background: Chromatography separates compounds based on the attraction between the compound and solvent (mobile phase) compared to that between the compound and the solid column (stationary phase). Artificial food dyes are complex organic molecules with a number of different functional groups (Schneider et al., 2007). The differences in the molecular structures affect the polarity of the molecules and, in turn, their attraction for the mobile phase versus the stationary phase. Dyes more attracted to the mobile phase will more quickly move through the column, whereas dyes more attracted to the stationary phase will better adsorb on the column and take longer to elute.

Activity: Use chromatography to separate the dyes in a grape drink such as Kool-Aid. Pretreat the Sep-Pak C18 column with 70% isopropyl alcohol and deionized water. Load the column by injecting it with grape drink from a 10-ml syringe. Note the color of the eluent. Repeat with 5% and 25% isopropyl alcohol noting the color of the eluent each time (University of Minnesota, 2000). Sep-Pak C18 cartridges, including instructions for this activity, may be purchased from Flinn Scientific.

Assessment: Ask students the following questions: (1) What serves as the stationary and mobile phases in this experiment? (2) Which eluant is the most polar? least polar? How do you know? (3) Which dye, red or blue, had the most attraction for the column? (4) Which dye had the strongest attraction for the isopropyl alcohol? (5) What is the purpose of using varying concentrations of isopropyl alcohol? How does this relate to the molecular structures of the dyes?

Extensions: Discuss health concerns of food dyes in the past. See *ChemCom* for an activity on the chromatography of food dyes in candy and Schneider et al. (2007) for a more advanced version of the *ChemCom* lab.

New problem: Most kidney stones are calcium stones (composed of calcium oxalate) and are formed according to the equation below:

$$\mathrm{Ca(HCO_3)_2\ (aq) + Na_2C_2O_4\ (aq) \rightarrow 2NaHCO_3(aq) + CaC_2O_4\ (s)}$$

Colas are high in oxalates and are not recommended for people who suffer from kidney stones. A can of cola contains (244 mg) of sodium oxalate, $Na_2C_2O_4$*. How many cans of soda would someone have to drink to produce a kidney stone with a mass of 1.00 g (1000 mg), assuming that there is plenty of calcium hydrogen carbonate available in the body?*

In *Living by Chemistry*'s toxins curriculum (Stacy et al., 2003*c*), students examine reaction types by looking at chemical changes that produce materials which, at certain concentrations, can be harmful to humans. Often, solution chemistry is introduced after chemical reactions; however, "Toxins" opens with an introduction to solutions and molarity to provide a context for the study of harmful chemicals. For students to track toxins, they must learn the language of chemistry (chemical equations). Later, students investigate precipitation reactions as the backdrop for stoichiometry and limiting reactants.

There are also some important biological reactions that involve smaller organic molecules that can be used to illustrate chemical reactions. One example is alcohol fermentation that is used by yeast and some bacteria ($C_6H_{12}O_6 \rightarrow 2C_2H_6O + 2CO_2$). As the products of this reaction are ethanol and carbon dioxide, humans have taken advantage of this reaction for years in the making of bread, beer, and wine. Students should easily be able to classify alcohol fermentation as a decomposition reaction.

Small organic molecules can be used in labs and stoichiometry problems as well. One good example that incorporates an organic acid with biological relevance is the making of fizzy drinks from citric acid and baking soda. In this activity, students work in the classroom or other location appropriate for food preparation (not the lab) and use stoichiometry to determine the correct proportion of ingredients (Rohrig, 2000). Students can easily evaluate their own results, as the incorrect ratio tastes bad. In addition to stoichiometry, this activity is ideal for equation writing and identifying the phases of reactants and products. Food-grade citric acid is inexpensive and may be purchased through online natural food outlets.

Thermodynamics

Thermochemistry is often taught after stoichiometry to help students understand how stoichiometric relationships dictate overall changes in energy during a chemical reaction. Some courses may delve deeper into thermodynamics later in the course. From qualitative descriptions of exothermic and endothermic reactions to quantitative descriptions of heat flow, entropy changes in reactions, and free energy changes used to predict reaction spontaneity, biological examples can help students make connections between biology and chemistry, as well as see the relevance of thermodynamics to their everyday lives. The most biologically significant process

Ready, Set, Break It Down Now

Purpose: Introduction to identifying chemical reactions

Background: Hydrogen peroxide (H_2O_2) is toxic to living organisms; therefore, it is quickly decomposed to yield H_2O and O_2 by the enzyme catalase. Although the decomposition of H_2O_2 is very slow under standard conditions; in one second, one molecule of catalase catalyzes the decomposition of 10^7 molecules of H_2O_2 (Moran et al., 1994).

Activity and assessment: Tell students that hydrogen peroxide (H_2O_2) is toxic to living organisms and ask them what compounds they think H_2O_2 could be broken down into that would be less toxic. If students say H_2 and O_2 gas, you can demonstrate how explosive this mixture is using a video clip from the internet (Purdue, 2002) or the *Journal of Chemistry Education* Chemistry Comes Alive video collection. Have students write a balanced chemical equation for the reaction they predict will occur. Add a small portion of fresh apple, fresh potato, or fresh raw liver to a test tube containing 5 ml of 3% H_2O_2 and describe what evidence indicates a chemical reaction is occurring (gas bubbles and test tube gets warm). You may want to have a brief discussion of catalysts at this point. You can also connect this activity to the elephant toothpaste demonstration (Spangler, 2007) and test the gas evolved with a glowing splint to confirm that it is oxygen.

Extensions: This can also be linked to kinetics and how catalysts or enzymes (nature's catalysts) speed up a reaction by lowering the activation energy. You can ask students what they might do to make the reaction go faster (for example, grind up the potato). Future labs can also include determining the optimal pH and temperature for enzymes to work. Example of such activities can be found in *ChemCom* (ACS, 2006).

that affects students is the metabolism of food. Metabolism can be examined on the macroscopic, cellular, and particulate levels. By showing students how chemical potential energy in foods can be converted into heat energy through combustion, they can appreciate the importance of understanding heat flow, calorimetry, and the relationship between heat and temperature in chemical changes. It is critical that experiments allow students to think about and discuss the difference between heat and temperature. Calorimetric experiments that allow students to calculate the energy released when food combusts can be used to help students relate joules to calories and kilocalories (Calories). Additionally, students can use the difference between energy released by food and energy gained by water (in a calorimeter) to explore what is meant by system, surroundings, and the First Law of Thermodynamics. Students should also come to understand that changes in energy do not always involve changes in temperature (e.g., phase changes). The standard exercise, in which students calculate the total number of calories required to change a known mass of ice at some temperature below 0°C to steam at some temperature above 100°C, could be made into a more interesting and biologically relevant problem by incorporating a weight loss myth. Because it takes energy to heat and/or melt cold water, ice, or even ice cream, some believe that drinking or eating these will lead to weight loss. The following problem helps students to distinguish between energy changes that involve temperature from those related to phase changes, highlights the difference between food Calories (kilocalories) and calories, and provides the "skinny" on this popular weight loss myth:

Suppose you eat two ice cubes (11.5 g) at –2.1 °C. Assuming that you don't choke or have a fever (your body temperature is 37.0°C), how much energy, in calories, is absorbed by the ice when it warms, melts, and warms to your body temperature? How many dietary Calories (kcal) were absorbed by the water? Is swallowing ice an effective weight loss method? Why or why not? Describe how the kinetic and potential energy of the ice change after you swallow it.

In addition to focusing on food to introduce the concept of calorimetry at the macroscopic level, revisiting the topic of cellular respiration from biology can link chemistry and biology, as well as focus on the particulate level, while addressing a common chemical misconception. Many students hold the belief that bond breaking is an exothermic process. A quick Internet search of high school biology Web sites reveals many statements describing how removing the third phosphate group of ATP releases a great deal of energy or free energy. The sites go on to define free energy as energy stored in chemical bonds. In addition, text and Web diagrams that depict the cleaving of the bond of one phosphate group from an ATP molecule often include an orange graphic labeled, "energy" appearing to be shooting out of the hydrolysis site. Although high school biology and chemistry courses are

The Heat Is On

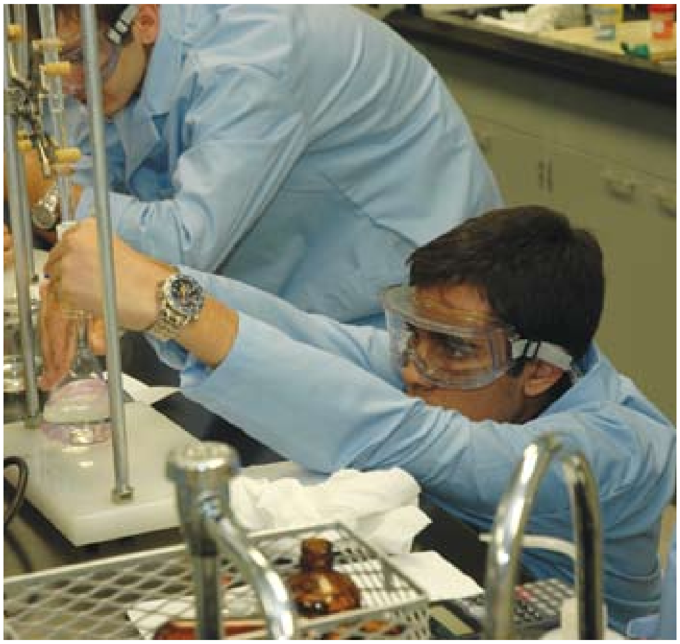
Purpose: To determine the heat content of hydrocarbons using calorimetry

Background: Burning of hydrocarbons in air (O_2) produces H_2O and CO_2. These products are lower in energy than the original hydrocarbons; thus energy is released in the form of heat. Similarly, our bodies break down food, releasing energy, which is reported as dietary Calories. In chemistry, we define a calorie as the energy needed to raise the temperature of 1 g of H_2O by 1°C. The dietary Calorie is equal to 1000 of these calories.

Activity: Students will use either candles or small oil lamps to determine the heat content of hydrocarbon fuels. To do this, students will need to know the initial and final masses of the heat source and the mass of the water, as well as the initial and final temperature of the water. There are several variations on the calorimeter; a popular model uses a soda can with a stirring rod threaded through the pop top suspended from a ring on a stand (ACS, 2006). Allowing the water temperature to rise 20°C will provide good results. Students should use their data and the heat capacity of H_2O to calculate the heat content of their fuel source. An alternative to using hydrocarbon fuels is to use a food source such as a nut, a marshmallow, or a cheeto. (Note that some schools place restrictions on the use of food because of students' allergic reactions.)

Assessment: Students can be asked the following questions: (1) Which fuel source has the highest heat content? Rank the remaining fuel sources from highest to lowest. (2) Write the balanced chemical equation for the combustion of wax ($C_{25}H_{52}$). (3) What mass of oil is needed to raise the temperature of 2 L of water from the room temperature to water's boiling temperature? (4) One Krispy Kreme original glazed doughnut contains 210 dietary Calories of chemical energy; how much wax (or oil) would you have to burn to release the same amount of energy as the doughnut?

Alternative: Students can burn a food source, determine the amount of energy released, and compare this to the nutritional information on the package. Cheetos and marshmallows are good alternatives to cashews for those concerned about nut allergies (*Journal of Chemical Education* Editorial Staff, 2004; Stacy et al., 2003a).



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responsible for the persistence of this misconception, this can be alleviated by drawing students' attention to the confusion surrounding this topic (Galley, 2004). A critique of resources showing the conversion of ATP to ADP may interest students and emphasize the notion that bond breaking is endothermic and bond making is exothermic. This could serve as an introduction to Hess's law, along with the sign conventions associated with enthalpies of formation (ΔH_f). As an extension, students may be interested in exploring macromolecules (fats, proteins, and carbohydrates) from a thermodynamic perspective. By using enthalpies of formation, students can determine why fat metabolism releases more calories of energy than the metabolism of proteins and carbohydrates.

Lastly, the Second Law of Thermodynamics can be introduced, but as it pertains to biological processes, such as photosynthesis and cellular respiration. Even a qualitative discussion of entropy is useful; however, be cautious, as defining a closed system around the complex cycle of photosynthesis and cellular respiration may be difficult. Students can summarize the net changes in enthalpy and entropy for photosynthesis and cellular respiration, with an emphasis on the energy changes that occur as a result of bond breaking and bond making and on the relative changes to the entropy of the system.

By recognizing where the physical and life science standards relevant to chemistry overlap, teachers can plan their courses to take advantage of students' experience and interest in the life sciences, thereby making chemistry accessible to more students.

Recommended Readings

ACS. *Chemistry in the Community*, 5th ed.; W. H. Freeman: New York, 2006.

A thematic based ACS textbook that includes units on Food; Energy Storage and Use; and Proteins, Enzymes, and Chemistry. Each chapter also contains various activities, some of which are highlighted in this chapter.

Stacy, A.M.; Coonrod, J.; Claesgens, J. *Living by Chemistry: General Chemistry. Unit 5 Fire. Teacher's Guide,* preliminary ed., Key Curriculum Press: Emeryville, CA, 2003*a*.

Stacy, A.M.; Coonrod, J.; Claesgens, J. *Living By Chemistry: General Chemistry. Unit 2 Smells. Teacher's Guide,* Key Curriculum Press: Emeryville, CA, 2003*b*.

Stacy, A.M.; Coonrod, J.; Claesgens, J. *Living By Chemistry: General Chemistry, Unit 4 Toxins. Teacher's Guide,* Key Curriculum Press: Emeryville, CA, 2003*c*. A full year high school chemistry curriculum that is divided into 6 thematic units. Each unit consists of 25–30 guided inquiry lessons, each 50 minutes in duration. The student and teacher materials have been field tested in high school classrooms. Biological chemistry examples from these materials have been highlighted throughout this chapter.

Recommended Web Sites

ACS Chem Club Student Resources. http://portal.acs.org/portal/acs/corg/content?_nfpb=true&_pageLabel=PP_ARTICLEMAIN&node_id=1505&use_sec=false (accessed March 11, 2008). An ACS Web site that has a number of links of interest to high school students. Under "What's that stuff" students can learn about materials such as lipstick, margarine, artificial sweeteners, food coloring, and more. There are also links for AAAS science podcasts, careers in chemistry, and reading up on chemistry topics.

McGill University Department of Chemistry. Office for Science and Society. http://oss.mcgill.ca/ (accessed March 11, 2008). A Web site highlighting chemistry in our everyday lives through its science issues and science in the news features. Additional features include a question-and-answer section and online lectures covering topics such as food chemistry, the diversity of chemistry, drugs, and chemistry and the environment. The lectures include visuals and audio.

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