CHAPTER 2: Atoms, Ions, and Molecules

**CHAPTER OVERVIEW**

This chapter is designed to help students understand the chemical level of organization. Atoms and molecules form the basis of human anatomical structures and physiologic processes, thus an understanding of this level of organization is very important. The chapter presents certain concepts in both inorganic chemistry and organic chemistry. It describes the chemical level of organization, all the way from the atom to the macromolecule.

The atom is discussed in terms of its construct. Various types of chemical bonds are introduced along with certain chemical properties associated with each type. The water molecule is presented along with water’s properties. An introduction to acids, bases, pH, and buffers is given in this chapter. Macromolecular structure and properties of carbohydrates, lipids, proteins, and nucleic acids are presented and discussed. All these chemical structures are presented in a confluent manner, allowing students to see interrelationships.

**KEY POINTS TO EMPHASIZE WHEN TEACHING BIOLOGICAL CHEMISTRY**

**An instructional understanding:** Most students taking an introductory course in human anatomy and physiology have difficulty understanding the chemical level of organization, particularly those who have not taken a good previous general biology or chemistry course. A basic understanding of atoms and their subatomic particles is important to an understanding of molecules and macromolecules found in the human body. Many students taking human anatomy and physiology are entering healthcare fields. It is important for these students and the others to understand the function and construct of the four basic macromolecules: carbohydrates, lipids, proteins, and nucleic acids. Some of the students will have to take additional courses in biology such as microbiology, nutrition, and pathophysiology; thus it is particularly important for them to obtain a firm foundation in biochemistry.

1. Explain to the student that **biological chemistry** is sometimes difficult to understand, unless the student has taken a previous course that has a chemistry component.
2. As often as possible, show practical applications of biological chemistry to the human body, particularly in respect to physiological processes and disease processes.
3. Define **chemistry** and how it relates to matter. For example, chemistry is the scientific study of matter.
4. Define **matter.**
5. Define and discuss **mass** and its relationship to matter.
6. Discuss how **weight** differs from mass.
7. Discuss the **atom** and how matter is related to it.
8. Define and discuss an **element** and how the atom relates to it.
9. Discuss the number of naturally occurring elements.
10. Discuss the major, lesser, and trace elements in the human body.
11. Show and discuss the **periodic table.**
12. Examine the structure and components of an atom.
13. Discuss the size, charge, and location of the **neutrons, protons,** and **electrons** in an atom.
14. Discuss **‘atomic mass unit’** and its relationship to a **‘Dalton’.**
15. Discuss the **chemical symbol.**
16. Explain the **atomic mass** and **atomic number.**
17. Explain how to determine the number of protons, neutrons, and electrons in an atom.
18. Show and explain the atomic structure in terms of the nucleus components (neutrons and protons) and the electron shells.
19. Define and discuss an **isotope.**
20. Discuss alterations in neutron number and atomic mass in an isotope.
21. Explain the isotope in terms of radiation.
22. Define and discuss a **radioisotope.**
23. Explain the **physical half-life** of a radioisotope.
24. Explain the **biological half-life** of a radioisotope.
25. Discuss the radioisotope’s importance in medical imaging.
26. Define and discuss an **ion.**
27. Explain some of the physiological roles of ions in the body.
28. Discuss ions in terms of electron number and charge.
29. Define a **cation** and **anion.**
30. Discuss **polyatomic ions.**
31. Discuss how to use the periodic chart to predict the charge on an ion.
32. Give examples of some of the ions in the body.
33. Explain how an ion relates to an electrolyte.
34. Define and discuss the **ionic chemical bond.**
35. Discuss why an ionic chemical bond can form between certain atoms.
36. Discuss the electron’s role in chemical bonding.
37. Discuss a **salt** in terms of ionic bonding.
38. Discuss **ionic compounds.**
39. Show and discuss types of chemical formulas.
40. Discuss and show examples of a **molecular formula.**
41. Discuss and show examples of a **structural formula** and how it differs from a molecular formula.
42. Explain how a molecule can have the same molecular formula but differ in the structural formula, thus forming an isotope.
43. Discuss the physiological importance of the various types of **isomers.**
44. Define and discuss the **covalent bond.**
45. Discuss why a covalent bond can form between certain atoms.
46. Explain the difference between an ionic bond and a covalent bond.
47. Explain the electron’s role in a covalent bond and how this role differs in an ionic bond.
48. Discuss the number **(single, double, triple) of covalent bonds** that can form between atoms and why.
49. Explain how carbon forms covalently bonded chains, termed the carbon skeleton.
50. Explain nonpolar covalent bonding and polar covalent bonding.
51. Discuss the properties and chemical behavior of polar covalent molecules and nonpolar covalent molecules.
52. Explain the structure of an **amphipathic molecule.**
53. Give examples of the physiologic roles of amphipathic molecules in the human body.
54. Define and the discuss **electronegativity,** in terms of chemical bonding.
55. Discuss intermolecular attractions.
56. Compare and contrast the ionic chemical bond and covalent bond in regards to intermolecular attractions.
57. Define and discuss **hydrogen bonding,** especially in terms of its importance to water.
58. Define and discuss **hydrophobic interactions.**
59. Define **van der Waals forces.**
60. Discuss water in terms of why is it so important to biological systems.
61. Show and discuss the molecular structure of water.
62. Repeat the discussion on the polar covalent bonding in water molecules, in terms of its importance to the chemical behavior of water.
63. Repeat the discussion on the hydrogen bonding between water molecules, in terms of its importance to the chemical behavior of water.
64. Discuss the importance of these two types of bonds in relation to the properties of water.
65. Discuss the liquid, gas, and solid phase of water and how water can transform between the three phases.
66. Discuss the transport, lubricant, cushioning, and excretion functions of water and how the two types of chemical bonds enable water to perform these functions.
67. Explain how hydrogen bonding results in **water’s adhesive** and **cohesive nature.**
68. Explain the physiological importance of water’s cohesive and adhesive nature.
69. Define temperature, specific heat, and heat of vaporization, and how it relates to water.
70. Discuss the reason for the tremendous solvent ability of water.
71. Discuss the types of substances that dissolve in water and the substances that do not dissolve, explaining the rationale for both actions.
72. Discuss the concept of a **“hydration shell”.**
73. Discuss and define dissociation of molecules in water.
74. Define **electrolytes** and **nonelectrolytes,** and how water relates to these substances.
75. Repeat the discussion of nonpolar molecules, but this time in reference to non-solubility in water.
76. Repeat the discussion of amphipathic molecules, but this time in terms of partial solubility in water.
77. Discuss acidic and basic solutions.
78. Define an **acid** and **base.**
79. Explain how the formation of an acid and base relates to the spontaneous dissociation of water.
80. Explain how acids and bases are formed.
81. Discuss pH in terms of its mathematical concept.
82. Discuss how pH measures acidity and basicity.
83. Show and discuss the pH scale in terms of acidic values, basic values, and neutral values.
84. Discuss the pH of certain foods and human body fluids.
85. Use outside sources to discuss the Henderson-Hasselbach pH equation used in biochemistry.
86. Define a **pH buffer** and how it is formed.
87. Discuss the role of pH buffers in the body.
88. Discuss the three types of **chemical mixtures.**
89. Define and discuss the properties of a **suspension.**
90. Define and discuss the properties of a **colloid.**
91. Define and discuss the properties of a **solution.**
92. Give examples of all three mixtures in the human body.
93. Discuss the homogenous nature of a solution.
94. Discuss measurements of solute concentrations.
95. Discuss the **mass/volume** measurement of solutes, giving examples like that of the concentration of glucose in the blood.
96. Define and discuss a **mole.**
97. Define and discuss **molecular mass.**
98. Discuss how the mole concept is important in the determination of solute concentrations.
99. Define and discuss **molarity.**
100. Define and discuss **molality.**
101. Compare and contrast molarity and molality.
102. Define and discuss an **osmole.**
103. Define and discuss **osmolarity.**
104. Define and discuss **osmolality.**
105. Compare and contrast osmolality and osmolarity.
106. Discuss general concepts concerning biological macromolecules.
107. Define **organic chemistry** and discuss how it differs from **inorganic chemistry.**
108. Repeat the discussion of why carbon is the basis of organic chemistry, again mentioning the carbon skeleton.
109. Explain the fact that all organic compounds contain carbon, but the compound must also contain hydrogen, thus a hydrocarbon. Carbon dioxide can be used as an example, since it contains carbon but not hydrogen, thus it is not organic.
110. Discuss certain commonalities among biological macromolecules such as the monomer to polymer relationship, dehydration synthesis, and hydrolysis decomposition.
111. Discuss the different types of functional groups found in biological macromolecules, explaining their respective chemical properties.
112. Discuss **lipids.**
113. Discuss the difference between a **saturated** fatty acidand an **unsaturated** one.
114. Define and show the structure of a triglyceride.
115. Discuss how **triglycerides** are formed in the human body.
116. Explain the anatomical and physiological functions of triglycerides in the human body.
117. Define and show the structure of a **phospholipid.**
118. Explain the physiological functions of phospholipids in the human body.
119. Define and show the structure of a **steroid.**
120. Explain the physiological functions of steroids in the human body.
121. Define and show the structure of an **eicosanoid.**
122. Explain the physiological functions of eicosanoids in the human body.
123. Discuss some of the other lipids, like **glycolipids** and the fat soluble vitamins.
124. Discuss **carbohydrates.**
125. Discuss the reason for the carbohydrate name: a hydrated carbon.
126. Define and show the structure of a **monosaccharide.**
127. Discuss the physiological importance of certain monosaccharides.
128. Discuss some of the **hexose monosaccharide isomers,** such as **glucose, fructose,** and **galactose.**
129. Discuss the two main pentose monosaccharides: ribose and deoxyribose.
130. Define and show the structure of a **disaccharide.**
131. Discuss the physiological importance of certain disaccharides, such as **lactose, sucrose,** and **maltose.**
132. Define and show the structure of a **polysaccharide.**
133. Discuss the physiological importance of certain polysaccharides, especially that of **glycogen.**
134. Discuss **glycogenesis** and **glycogenolysis.**
135. Discuss the nucleic acids.
136. Define a **nucleic acid.**
137. Define and show the composition and structure of the **monomeric unit** of the nucleic acids, the nucleotide.
138. Discuss the structure of the two general classes of nitrogenous bases: **purine** and **pyrimidine.**
139. Discuss and show the structure of the two purine nitrogenous bases: **adenine** and **guanine.**
140. Discuss and show the structure of the three pyrimidine nitrogenous bases: **cytosine, uracil,** and **thymine.**
141. Discuss and show how nucleotides covalently bond to form the nucleic acid polymers: **deoxyribonucleic acid (DNA)** and **ribonucleic acid (RNA).**
142. Discuss and show the structuring of DNA.
143. Discuss and show the structuring of RNA.
144. Discuss the function and show the structure of another nucleotide, **adenosine triphosphate,** termed **ATP.**
145. Discuss the function and show the structures of the nucleotides: **nicotinamide adenine dinucleotide (NAD)** and **flavin adenine dinucleotide (FAD).**
146. Discuss the proteins.
147. Define a **protein.**
148. Discuss seven functions of proteins: enzyme, defense, transport, support, movement, regulation, and storage.
149. Define and show the structure of the monomeric unit of a protein, the amino acid.
150. Discuss and show the structure of the 20 naturally occurring amino acids.
151. Explain the role of the carboxyl group in an amino acid.
152. Explain the role of the amine group in an amino acid.
153. Explain that the differences in both structure and chemical behavior among the amino acids are based on the types of R groups.
154. Discuss and show structures of nonpolar amino acids, polar amino acids, charged amino acids, and amino acids with special functions.
155. Explain how amino acids use dehydration synthesis to form covalent bonds, known as peptide bonds, to form a polymer.
156. Define the various polymer terms used when amino acids covalently bond together: dipeptide, oligopeptide, polypeptide, and protein.
157. Define the term **glycoprotein.**
158. Discuss the fact that protein molecular shape is paramount to its functioning.
159. Discuss the conditions that cause a protein to bend out of shape, termed **denaturation.**
160. Define and discuss the **primary structure** of a protein.
161. Discuss how the primary structure, linear sequence of amino acids, determines the three- dimensional **conformation** of a protein.
162. Define and discuss the **secondary structure** of a protein, reiterating how the primary structure is primarily responsible for determining the secondary structure.
163. Define and discuss the **tertiary structure** of a protein, reiterating how the primary structure is primarily responsible for determining the tertiary structure.
164. Define and discuss the **quaternary structure** of a protein, reiterating how the primary structure is primarily responsible for determining the quaternary structure.
165. Discuss the role of chaperone proteins in the folding of a protein.

**ADDITIONAL TOPICS FOR DISCUSSION**

1. Explain and discuss the various types of formulas and models: empirical formula, molecular formula, structural formula, and molecular models.
2. Explain and discuss optical isomerism, in terms of cis and trans, noting its importance in nutrition.
3. Explain and discuss the types of covalent bonding.
4. Discuss why carbon is the basic element in biological systems.
5. Discuss the relative strengths of acids and bases.
6. Discuss pH using the Henderson-Hesselbach equation.
7. Discuss the nutritional and physiological importance of certain plant polysaccharides, such as cellulose and amylose.
8. Discuss the structure of various fatty acids.
9. Discuss the nutritional and physiologic importance of the omega fatty acids.
10. Discuss some of the problems encountered when there are errors in the primary structure of a protein.

**SUGGESTED CHAPTER OUTLINE**

**2.1 Atomic Structure: The simplest level of organization is composed of atoms, ions, and molecules. (pp. 32–35)**

A. Matter, Atoms, Elements, and the Periodic Table (p. 33)

1. The human body is composed of matter, defined as a substance that has mass and occupies space.

2. Matter is present in the human body in three forms: solid, liquid, and gas.

3. All matter is composed of atoms.

4. An atom is the smallest particle that exhibits the chemical properties of an element.

5. There are 92 naturally occurring elements, with hydrogen being the smallest and uranium the largest and heaviest.

6. Technical advances in chemistry and physics have resulted in the ability to produce “ultraheavy” elements that are larger than uranium.

7. All elements are organized into a chart form in the periodic table of elements.

8. Elements are grouped into major, lesser, and trace elements based on the percentage each composes by weight in the human body.

9. Major elements comprise 98% and minor elements less than 1%.

10. Only 12 elements occur in living organisms in greater than trace amounts: oxygen, carbon, hydrogen, nitrogen, calcium, phosphorus, sulfur, potassium, sodium, chlorine, magnesium, and iron.

11. Atoms are composed of three subatomic particles: protons, neutrons, and electrons.

12. Two major criteria differentiate subatomic particles—namely, mass and charge.

13. The mass of an atom is expressed as the atomic mass unit (amu).

14. Neutrons are uncharged, having a mass of 1 amu.

15. Protons have a positive charge, also having a mass of 1 amu.

16. Neutrons and protons are located in the atomic nucleus.

17. The electrons have a negative charge, having a mass of 1/800th the mass of a proton or neutron.

18. Electrons are located at varying distances from the nucleus in regions called orbitals, often depicted as either an electron cloud or discrete energy levels.

19. Elements differ in subatomic particles, and the periodic table can be used to determine the number of these subatomic particles.

20. The periodic table shows an element’s symbol, atomic number, and average atomic mass.

21. A unique chemical symbol is assigned to each element.

22. The atomic number of an element indicates the number of protons in an atom.

23. The average atomic mass indicates the mass of both protons and neutrons in the atomic nucleus, and it reflects the “heaviness” of an element’s atoms relative to atoms of other elements.

24. The number of protons in an atom is the atomic number.

25. The number of neutrons can be determined by subtracting the atomic number from the atomic mass.

26. The number of electrons is the same as the atomic number, because all atoms are neutral.

27. Shells of electrons surround the atomic nucleus, and each shell has a given energy level.

28. Each shell can only hold a limited number of electrons, with the innermost shell holding up to two electrons, and the second shell holding up to eight electrons.

29. All subsequent shells also house eight electrons, but some subsequent shells hold more than eight.

B. Isotopes (p. 34)

1. Isotopes are different atoms of the same element that have the same number of protons and electrons but differ in the number of neutrons.

2. Isotopes exhibit essentially identical chemical characteristics but have a different atomic mass.

3. Carbon exists in three isotopes: carbon-12, carbon-13, and carbon-14.

4. All the carbon isotopes have six protons in their nuclei; however, carbon-12 has six neutrons, carbon-13 has seven neutrons, and carbon-14 has eight neutrons.

5. Some isotopes are referred to as radioisotopes, in that they are unstable due to having an excess number of neutrons.

6. Radioisotopes usually lose nuclear components in the form of high-energy radiation that includes alpha particles, beta particles, or gamma rays as they decay or break down to a more stable isotope.

7. The time it takes for 50% of the radioisotope to become stable is its physical half-life.

8. The time it takes for half of the radioactive material (e.g., from a medical test using radioactive contrast material) to be eliminated from the body is the biological half-life.

C. Chemical Stability and the Octet Rule (pp. 34−35)

1. The periodic table is organized into rows based on atomic number, and it is organized into columns based on electrons in the outer shell, or what is referred to as the valence shell.

2. Each consecutive column, IA to VIIIA, has the same number of electrons in its valence shell, as represented by the numeral at the top of the column.

3. For example, in column IIA, all the elements in the column have two electrons in their valence shells.

4. Atoms that have a completely filled valence shell of electrons are stable, and thus do not combine with other elements.

5. Elements that do not have a filled valence shell tend to lose, gain, or share electrons to obtain a complete outer shell; they are termed inert.

6. The tendency to fill the valence shell is termed the octet rule.

**2.2 Ions and Ionic Compounds: The body is composed mostly of compounds. (pp. 36–38)**

A. Ions (pp. 36–37)

1. Ions are atoms or groups of atoms with either a positive charge or a negative charge and are produced from the loss or gain of an electron or electrons.

2. The decision whether an atom will lose or gain electrons depends on which action will provide it with a complete valence shell.

3. When an atom loses one or more electrons, it becomes positively charged.

4. When at atom gains one or more electrons, it becomes negatively charged.

5. Ions with a positive charge are called cations, and ions that are negatively charged are termed anions.

6. When two or more atoms complex together and become ions, they are termed polyatomic ions.

7. The periodic chart can be used to predict whether an atom will become a cation or an anion.

8. Usually, elements on the left side of the periodic chart lose electrons, thus becoming cations, and elements on the right side of the chart gain electrons, thus becoming anions.

B. Ionic Bonds (pp. 37−38)

1. Positively charged cations and negatively charged anions may bind together by electrostatic interactions called ionic bonds.

2. The structures formed are salts.

3. A classic example is when table salt is formed from the bonding of metallic atoms of sodium with nonmetallic atoms of chlorine.

4. In this example, the sodium atom donates one outer shell electron to a chlorine atom.

5. After donating the electron, the sodium atom becomes a cation (Na+) and the chlorine atom becomes an anion (Cl−).

**2.3 Covalent Bonding, Molecules, and Molecular Compounds: Instead of donating or gaining electrons to form ionic compounds, atoms also have the possibility of reaching stability by sharing electrons. (pp. 38–42)**

A. Chemical Formulas: Molecular and Structural (p. 38)

1. The sharing of electrons between atoms results in a covalently bonded molecule.

2. Most molecules are composed of two or more different elements and are called molecular compounds.

3. The chemical constituents of a molecule and their ratios are represented using a molecular formula.

4. The structural formula of a molecule is complementary to its molecular formula and exhibits not only the number and types of atoms but also their arrangements within the molecule.

5. Structural formulas provide a means for differentiating isomers, which are molecules composed of the same number and kind of elements but arranged differently in space.

6. Isomers may have very different properties from one another—therefore, structural formulas are an important piece of chemical information.

B. Covalent Bonds (pp. 39–41)

1. The bond that is formed when atoms share electrons is a covalent bond.

2. Covalent bonding occurs when both atoms require electrons to become stable.

3. This bonding takes place when the participating atoms that form the chemical bond have four, five, six, or seven electrons in the outer shell.

4. The most common elements of the human body that form covalent bonds are oxygen, carbon, hydrogen, and nitrogen.

5. The number of covalent bonds formed by an atom may be determined by examining the number of electrons needed to complete the outer shell.

6. Atoms of elements can form more than one covalent bond also may do so through combinations of single, double, or triple covalent bonds.

7. A single covalent bond is one pair of electrons shared between two atoms.

8. A double covalent bond is two pairs of electrons shared between two atoms.

9. A triple covalent bond is three pairs of electrons shared between two atoms.

10. Numerous carbon atoms are sometimes bonded together to form a 'carbon skeleton'.

11. Three possible arrangements of the carbon skeleton may occur: a straight chain, branched chain, or a ring.

12. Atoms share electrons in a covalent bond either equally or unequally between the atoms.

13. How the atoms share the electrons between them is determined by the relative attraction each atom has for electrons, a concept referred to as electronegativity.

14. Different types of atoms have varying degrees of electronegativity, or attraction for electrons, and thus may share the electrons unequally, resulting in what is termed a polar covalent bond.

15. If the atoms share the electrons equally, a nonpolar covalent bond is formed.

C. Nonpolar, Polar, and Amphipathic Molecules (pp. 41−42)

1. If the atoms share the electrons equally, a nonpolar covalent bond is formed.

2. Nonpolar molecules in the body are formed predominantly by covalent bonding between the same elements.

3. Polar molecules are formed by covalent bonding between different elements.

4. Sometimes a molecule is large enough that it can have one major part that is nonpolar and another part that is polar, resulting in a molecule termed an amphipathic molecule.

D. Intermolecular Attractions (p. 42)

1. Molecules sometimes have weak chemical attractions to other molecules, called intermolecular attractions.

2. One important intermolecular attraction is termed a hydrogen bond, which occurs between polar molecules.

3. A hydrogen bond is a weak attraction between a partially positive hydrogen atom within a molecule and a partially negative atom within another molecule.

4. Another type of intermolecular attraction is the van der Waals forces.

5. Van der Waals forces occur when electrons orbiting the nucleus of an atom of a nonpolar molecule are for a brief instant distributed unequally, causing one portion of the atom to be slightly negative and one end slightly positive.

6. Nonpolar atoms with an unequal charge distribution cause neighboring atoms to perform similarly, leading to opposition of charges that cause weak intermolecular attractions:   
van der Waals forces.

7. Hydrophobic interactions are a type of intermolecular attraction between nonpolar molecules when placed in water.

**2.4 Molecular Structure of Water and the Properties of Water: Water is the substance that comprises approximately two-thirds of the human body by weight. (pp. 43–46)**

A. Molecular Structure of Water (p. 43)

1. Chemist classify molecules into two broad categories: organic molecules and inorganic molecules.

2. Organic molecules are defined as molecules that contain carbon, which are (or have been) components of living organisms, such as glucose, protein, and triglycerides.

3. Inorganic molecules are all the other molecules that are not organic, such as water, salts, and others.

4. Water is a polar molecule composed of one oxygen atom bonded to two hydrogen atoms (H2O).

5. The polar nature of water is due to the oxygen atom being more electronegative than the hydrogen atom, thus pulling the electrons unequally.

6. Every water molecule has the ability to form four hydrogen bonds with adjacent water molecules.

B. Properties of Water (p. 44)

1. Water is present in three phases, depending upon the temperature: a gas (water vapor), a liquid (water), and a solid (ice).

2. Water has several functions in the human body: transport, lubricate, cushion, and excrete wastes.

3. Cohesion is the attraction between water molecules, due to hydrogen bonding between the water molecules.

4. Surface tension is the inward pulling of cohesive forces at the surface of water, due to water at the surface only being able to form three hydrogen bonds rather than the four formed by water in the internal liquid.

5. Adhesion is the attraction between water molecules and a substance other than water, as a result of water forming hydrogen bonds with molecules of the other substance.

6. Temperature is a measure of the kinetic energy, or random movement, of atoms or molecules within a substance.

7. The relationship between temperature and kinetic energy is direct—the temperature is higher when the kinetic energy is greater.

8. Specific heat is the amount of energy required to increase the temperature of 1 gram of a substance 1 degree Celsius.

9. Water has a very high specific heat as a result of the need to break hydrogen bonds in water to raise its temperature.

10. Heat of vaporization is the heat required for the release of molecules from a liquid phase into a gaseous phase for 1 gram of a substance.

11. Water has a high heat of vaporization as a result of the need to break hydrogen bonds to release water from the liquid into the gaseous phase.

C. Water as the Universal Solvent (pp. 44−46)

1. Water is the solvent of the body, and substances that dissolve in water are called solutes.

2. Many substances dissolve in water, but not all.

3. The chemical properties of a substance (whether it is polar, charged, nonpolar, or amphipathic) determine how it interacts with water.

4. Since water is charged (polar), substances that can dissolve in it are also charged, such as ions and polar molecules.

5. Substances that can dissolve in water are called hydrophilic.

6. Substances that dissolve in water are surrounded by many water molecules, known as a hydration shell.

7. Some substances dissolve in water but do not remain intact, thus they break apart: dissociate.

8. Dissociation in water occurs with substances that have ionic bonding.

9. Acids and bases dissociate in water.

10. Substances that both dissolve and dissociate in water, such as salts, acids, and bases, can readily conduct an electric current, thus they are called electrolytes.

11. Hydrophobic substances do not dissolve in water, due to these substances not being charged: nonpolar.

12. Nonpolar substances are excluded from water, thus they are hydrophobic.

13. The interaction between the molecules of the “excluded” nonpolar excluded substance is termed hydrophobic interaction because it appears these molecules are avoiding water.

14. Amphipathic molecules have both polar and nonpolar regions.

15. Amphipathic molecules partially dissolve in water, due to the polar region seeking water and the nonpolar region avoiding water.

16. Certain molecules (phospholipids) in the cell membrane are amphipathic.

**2.5 Acidic and Basic Solutions, pH, and Buffers: Acidic and basic solutions occur when an acid or base is added to water (pp. 46–49)**

A. Water: A Neutral Solvent (pp. 46–48)

1. Water can spontaneously depolarize as a result of the covalent chemical bond between oxygen and either of the two hydrogen atoms in a water molecule spontaneously breaking apart at a low rate, about 10−7 ions per liter.

2. When a hydrogen ion transfers to a water molecule, giving it an extra hydrogen ion (H+), it is represented as H30+, and termed the hydronium ion.

3. When a water molecule loses a hydrogen ion, it is termed a hydroxide ion, and represented as OH−.

4. Water is neutral in that it has as many positively charged hydronium ions as it does negatively charged hydroxide ions.

B. Acids and Bases (p. 48)

1. An acid is a substance that dissociates in water to produce an H+ and an anion, thus it is also called a proton donor.

2. A base accepts H+ when added to a solution, thus it is termed a proton acceptor.

C. pH, Neutralization, and the Action of Buffers (pp. 48–49)

1. The pH of a solution is a measure of the relative amounts of H+, and is expressed as a number between 0 and 14.

2. The pH is calculated by taking the inverse logarithmic value of the hydrogen ion concentration [H+].

3. The “inverse” logarithmic relationship causes the scale to appear inverted, in that the higher the hydrogen ion concentration, the lower the number, and the lower the hydrogen ion concentration, the higher the number.

4. Pure water and other solutions that have equal concentrations of H+ and OH− are neutral and have a pH of 7.

5. Solutions with a pH below 7 are acidic, and solutions with a pH above 7 are basic, or alkaline.

6. Neutralization occurs when a solution that is either acidic or basic is returned to neutral (pH 7).

7. A buffer is a single substance, or an associated group of substances, that function to help prevent pH changes if either excess acid or base is added.

**2.6 Water Mixtures: Mixtures are formed from the combining, or “mixing” of two or more substances. (pp. 49–51)**

A. Categories of Water Mixtures (pp. 49–50)

1. Water mixtures are placed into three categories based on the relative size of the substance mixed with water and include suspensions, colloids, and solutions.

2. A suspension, like sand in water, is formed when a material that is larger than 100 nanometers particulate size is mixed with water, resulting in a mixture that does not remain mixed unless it remains in motion.

3. Blood cells within the plasma (the liquid portion) form a suspension.

4. A colloid is a mixture composed of protein within water, where the protein ranges in size from 1 to 100 nanometers.

5. Colloids in the body include semifluid cell cytosol and proteins within plasma.

6. A solution is a homogeneous mixture in which the particulate size of the substance is smaller than 1 nanometer, and it dissolves in water.

7. An emulsion is a type of suspension composed specifically of water and a nonpolar (hydrophobic) liquid substance such as vegetable oil in water; breast milk is an example of an emulsion.

B. Expressions of Solute Concentration (pp. 50–51)

1. The amount of solute dissolved in a solution determines the concentration of a solution, and may be expressed in several ways: mass/volume, mass/volume percent, molarity, and molality.

2. Mass/volume is grams of solute per volume of solution.

3. Mass/volume percent is grams of solute per 100 milliliters of solution.

4. Molarity is a measure of number of moles per liter of solution.

5. Molality is the moles per kilogram of solvent.

6. An osmole (osm) is another means of expressing concentration, which reflects whether a substance either dissolves, or dissolves and dissociates, when placed in a solution (i.e., whether it is a nonelectrolyte or electrolyte).

7. An osmole measures the number of osmotic active particles in a solution.

8. Osmoles can be expressed as either osmolarity or osmolality.

9. Osmolarity is the number of particles in 1 liter of solution, whereas osmolality is the number of particles in 1 kilogram of water.

10. A mole is the mass in grams that is equal to either the atomic mass of an element or the molecular mass of a compound.

11. One mole contains 6.02 x 1023 atoms, ions, or molecules.

**2.7 Biological Macromolecules: Four classes of organic biological macromolecules (biomolecules) can be distinguished in living organisms: lipids, carbohydrates, nucleic acids, and proteins. (pp. 51–63)**

A. General Characteristics (pp. 51–53)

1. Biological macromolecules are large organic molecules that are synthesized by the

human body.

2. Biological macromolecules always contain carbon and hydrogen, and generally oxygen; some also contain nitrogen, phosphorus, and sulfur.

3. Organic biological macromolecules have several distinct features: carbon and carbon skeletons, hydrocarbons, and functional groups.

4. Carbon is the central element of biological macromolecules, existing independently, or covalently bonded into a carbon skeleton.

5. Hydrocarbons contain both carbon and hydrogen, covalently bonded together.

6. Functional groups are two or more atoms bonded together, attached to a carbon skeleton, displaying the same specific chemical characteristics, no matter what carbon skeleton they are attached to.

7. Polymers, present in biological macromolecules, are composed of repeating subunits called monomers.

8. Polymer forms are seen in carbohydrates, proteins, and nucleic acids, but not in lipids.

9. Dehydration synthesis is a process that combines monomers to form polymers.

10. Dehydration synthesis involves the removal of a —H from one monomer and an —OH from another monomer, thus allowing the formation of a covalent bond between the two monomers, and resulting in the formation of a water molecule when the —H bonds to

the —OH (H2O).

11. Hydrolysis decomposition is a process that splits polymers into monomers by inserting a split water molecule into the polymer, thus breaking the covalent bonding between the two monomers.

12. In the process of hydrolysis decomposition, water is split into a —H and —OH, the —H is added to one monomeric subunit and the —OH is added to the other subunit, resulting in a split covalent bond between the two monomers.

B. Lipids (pp. 53–56)

1. Lipids are a diverse group of macromolecules that do not form polymers, and are totally (nonpolar) or partially (amphipathic) insoluble in water.

2. Lipids function as stored nutrients, components of cell membranes, and hormones.

3. The four primary classes of lipids are triglycerides (neutral fats), phospholipids, steroids, and eicosanoids.

4. Triglycerides are formed from a glycerol molecule and three fatty acids bonded together.

5. Triglycerides are the most common form of lipids in living things and function to store energy, and provide structural support, cushioning, and insulation of the body.

6. A fatty acid is a lipid molecule with generally 14 to 20 carbons bonded together in a chain, and terminating in an acidic carboxyl functional group.

7. If the fatty acid molecule contains a double bond it is termed unsaturated, and if it does not contain a double bond, it is termed saturated.

8. If the fatty acid has one double bond it is unsaturated and if it has more than one it is polyunsaturated.

9. Triglycerides are a storage form of energy located in fat (adipose) cells.

10. When cells form triglycerides for energy storage, it is termed lipogenesis, and when cells break down triglycerides to release energy, it is termed lipolysis.

11. A phospholipid is amphipathic, and has a chemical structure similar to a triglyceride, except the phospholipid molecule contains a polar phosphate functional group.

12. Phospholipids are amphipathic in that they contain a polar head, where the polar phosphate group is positioned, along with nonpolar tails, where the fatty acids are positioned.

13. Phospholipids are a main component of cell membranes.

14. Steroids are composed predominantly of hydrocarbons arranged in a distinct four multi-ringed structure, three of the rings having six carbons and one ring having five carbons.

15. Steroids differ according to the side chains extending from the rings.

16. Steroids include cholesterol, steroid hormones (e.g., testosterone, estrogen, and progesterone), and bile salts.

17. Eicosanoids are modified 20-carbon fatty acids that are synthesized from arachidonic fatty acid, a common component in cell and nuclear membranes.

18. Four classes of eicosanoids are produced and include prostaglandins, thromboxanes, leukotrienes, and prostacyclins.

19. Glycolipids are lipid molecules with a carbohydrate covalently attached, having functions associated with cell membranes, such as cell binding.

20. Vitamins A, E, and K are lipids known as the fat soluble vitamins.

C. Carbohydrates (pp. 56–57)

1. The term carbohydrate is derived from the fact that a monomer of carbohydrate, termed a monosaccharide, has virtually every carbon hydrated with the equivalent of a water molecule, with an —H and —OH usually attached to each carbon.

2. One monomer of carbohydrate is termed a monosaccharide, two covalently bonded together into one molecule is termed a disaccharide, and several bonded together is a polysaccharide.

3. Glucose, fructose and galactose are six-carbon monosaccharide isomers, termed hexose sugars.

4. Ribose and deoxyribose are five-carbon monosaccharides, termed pentose sugars, and are found in RNA (ribonucleic acid) and DNA (deoxyribonucleic acid), respectively.

5. Sucrose (table sugar), lactose, and maltose are common disaccharides.

6. Glycogen, the storage carbohydrate in animals, is a polysaccharide, while cellulose (fiber) is a plant polysaccharide.

D. Nucleic Acids (pp. 57–59)

1. Nucleic acids are macromolecules that store and transfer genetic information, and were initially discovered in the cell nucleus.

2. Two classes of nucleic acid are deoxyribonucleic acid (DNA) and ribonucleic acid (RNA).

3. A nucleic acid monomer, known as a nucleotide, is composed of a nitrogenous base, pentose sugar, and phosphate group, covalently bonded together.

4. The pentose sugar in DNA is deoxyribose, and the pentose sugar in RNA is ribose.

5. There are five nitrogenous bases classified as either pyrimidines, which are single-ringed, or purines, which are double-ringed.

6. The nitrogenous bases adenine and thymine are pyrimidines, and cytosine, uracil, and guanine are purines.

7. Deoxyribonucleic acid (DNA) is a double-stranded nucleic acid; it can be found both as components of chromosomes within the cell nucleus and as a circular strand of DNA found in the mitochondria.

8. The nitrogenous bases in DNA are adenine, cytosine, guanine, and thymine; DNA contains no uracil.

9. The nitrogenous bases in RNA are adenine, cytosine, guanine, and uracil; RNA contains no thymine, it is replaced by uracil.

10. RNA is a single-stranded nucleic acid located both within the cell nucleus and within the cytoplasm of the cell.

11. Adenosine triphosphate (ATP) is another nucleotide that is composed of adenine, ribose, and three phosphate groups; it provides chemical energy within the cell.

12. Nicotinamide adenine dinucleotide (NAD+) and flavin adenine dinucleotide (FAD) are nucleotides that participate in the production of ATP in the cell mitochondria.

E. Proteins (pp. 59–63)

1. Proteins are polymers composed of one or more linear strands of amino acid monomers that may number in the thousands.

2. Proteins have many functions: enzyme, defense, transport, support, movement, regulation, and storage.

3. Proteins are composed of twenty different amino acids.

4. An amino acid, the monomeric unit of a protein, is a molecule composed of an acidic carboxyl function group, and a basic amine group.

5. The differences between amino acids are due to side-chain structures, termed the R groups.

6. Amino acid chemical properties are determined by the chemical nature of the R group; some amino acids are nonpolar, some polar, some charged, and some have special functions.

7. Amino acids are covalently linked together by a peptide bond, which is a type of covalent bond, found only in amino acid polymers, and formed by dehydration synthesis.

8. Two amino acids bonded together into one molecule is termed a dipeptide, 3 to 20 bonded together is termed an oligopeptide, 21 to 199 bonded together is termed a polypeptide, and 200 or more is termed a protein.

9. A protein covalently bonded to carbohydrate is termed a glycoprotein; the ABO surface marker on a red blood cell is one example of glycoprotein function.

**2.8 Protein Structure: Protein structure is paramount to its functioning. (pp. 64–67)**

A. Categories of Amino Acids (pp. 64–66)

1. Amino acids are organized into four groups based on the chemical characteristics of their R group: nonpolar amino acids, polar amino acids, charged amino acids, and amino acids with special functions.

2. Nonpolar amino acids contain R groups with either hydrogen (glycine) or hydrocarbons (alanine, valine, leucine, isoleucine, phenylalanine, and tryptophan); they tend to group with other nonpolar amino acids by hydrophobic interactions in water.

3. Polar amino acids contain R groups with elements in addition to carbon and hydrogen (e.g., O, N, or S) (serine, threonine, asparagine, glutamine, and tyrosine); they form interactions with other polar molecules and with water.

4. Charged amino acids can either have a negative charge or a positive charge.

5. Negatively charged amino acids, with a negatively charged R group, include glutamic acid and aspartic acid, and those with a positively charged R group include histidine, lysine, and arginine.

6. Charged and polar amino acids are hydrophilic, and their presence increases the solubility of the protein in water.

7. Three amino acids have unique characteristics: proline, cysteine, and methionine.

8. Proline has an R group that attaches to the amino group, forming a ring that bends proteins.

9. Cysteine is an amino acid containing a sulfhydryl group, allowing it to form a disulfide covalent bond (bridge) between the sulfhydryl group of another cysteine amino acid; these disulfide bridges stabilize the construct of a protein.

10. Methionine is always the first amino acid positioned when a protein is synthesized.

B. Amino Acid Sequence and Protein Conformation (pp. 66–67)

1. The structuring of a protein is described on a hierarchical level, having four possible structural levels: primary, secondary, tertiary and quaternary.

2. The primary structure is the linear sequence of amino acids in the protein.

3. The next three structural levels are responsible for the three-dimensional shape of a protein, known as protein conformation.

4. The more complex structural organizations of a protein are dependent upon intramolecular attractions between the amino acids in the linear sequence (primary structure) for proper folding and maintaining of a protein’s conformation.

5. Proteins, known as chaperone proteins, assist in proper protein folding.

6. Four main types of intermolecular and intramolecular interactions contribute to the final conformation of a protein: hydrophobic exclusion, hydrogen bonding, ionic bonding, and disulfide bonds.

7. The primary protein polymer is forced into its initial shape as hydrophobic exclusions “tuck” amino acids with nonpolar R groups into a more central location, limiting their contact with water.

8. Hydrogen bonds form between polar R groups, and between amine and carboxylic acid functional groups of closely positioned amino acids.

9. Ionic bonds form between negatively charged and positively charged R groups.

10. Disulfide bonds form between the sulfhydryl groups of two cysteine amino acids.

11. The secondary structure of a protein is a series of repeating patterns within the protein.

12. There are two possible secondary repeating patterns: alpha helix and beta-pleated sheet.

13. An alpha helix pattern is a spiral coiling arrangement of the protein, whereas the beta-pleated sheet is a planar arrangement.

14. The alpha helix gives some elasticity to fibrous proteins that are located in skin or hair.

15. The beta-pleated sheet gives some degree of flexibility to many globular proteins.

16. The tertiary structure of a protein is the final three-dimensional shape exhibited by a completed polypeptide chain.

17. Two categories of proteins, either fibrous or globular, are distinguished by their molecular shape.

18. Globular proteins fold into a compact, often nearly spherical shape such as enzymes, antibodies, and some hormones.

19. Fibrous proteins are extended linear molecules that are constituents of ligaments, tendons, and contractile proteins within muscle cells.

20. The quaternary structure of a protein is present only in those proteins with two or more polypeptide chains; hemoglobin is an example of a protein in the quaternary structure.

21. The normal function of a protein may also require a prosthetic group which is a nonprotein structure covalently bonded to a protein; heme in hemoglobin is an example of a prosthetic group.

22. The biological activity of a protein is usually disturbed or terminated when its conformation is changed, termed denaturation.

23. Denaturation occurs when a protein is subjected to a non-optimal chemical environment, such as improper pH and improper temperature.

**Visuals, In-Class Demonstrations, and Discussions**

1. Using a molecular modeling kit, demonstrate the manipulation of atoms with bonds.
2. Using the VSPER Theory, have students assign geometry, molecular structure, and hybridization to five substances.
3. Using a dart, demonstrate where electrons will fall outside the nucleus. If darts are not an option in your classroom, use the split pea orbitals model.
4. Show the flame test of seven elements and have the students complete the Abigail Freiberger’s Flame test demonstration worksheet.
5. Using the SMART Board, demonstrate how images produced by computed tomography (CT) and magnetic resonance imaging (MRI) can be used to estimate an individual’s body fat content.
6. One of the problems with discussing the atom is that we cannot see it. To help students overcome this, use nuts and bolts as models to visualize and explain the behavior of atoms.
7. Using the SMART Board, present PowerPoint slides on atoms and molecules, and how they form the basis of human anatomical structures and physiological processes.
8. Show the periodic table and discuss the trends of the periodic table.
9. Explain how molecular and structural formulas symbolize the composition of compounds. Discuss the size, charge, and location of the neutrons, protons, and electrons in an atom.
10. Using milk of magnesia, demonstrate acid and bases. Discuss the changes in pH.
11. Demonstrate osmosis by allowing students to observe the effects of different concentrations of salt solutions on potato cores.
12. Demonstrate how polypeptides may become folded through interactions between the side chains of amino acids.
13. Chart the major groups of inorganic chemicals common in cells and chart the functions of each of the chemicals.

**Related Media**

*Physical Science in Action: Atoms and Molecules.* Museum of Science.

*Diffusion and Osmosis Instructional.* DVD, Carolina.

*Overview of the Periodic Table.* School Media Associates.

*Chemical Reaction.* Museum of Science.

*In Matter: Atoms, Elements, & Chemistry. Squibs Vol.5,* Museum of Science.

*Atom.* Ambrose Video.

*The World of Chemistry.* The University of Maryland and the Educational Film Center.

*A Matter of State.* Museum of Science.

*Covalent Bonding.* File Transit.

*Chemical Bonds.* Museum of Science.

*Chemistry: Lesson 27: Strong Acids and Bases 1.* TMW Media Group.

*Water.* Museum of Science.

*Atoms, Bonding, and Structure.* 2.0 File Transit.

*The Proton in Chemistry.* Museum of Science.

*Ionic Bonding Software.* File Transit.

*Atoms and Their Electrons.* Clearvue and SVE, Inc.

*Bonding Between Atoms.* Clearvue and SVE, Inc.

*Chemical Reactions.* Clearvue and SVE, Inc.

*Atomic Number, Mass Number and Isotopes Part 1.* TMW Media Group.

*Atomic Number, Mass Number and Isotopes Part 2.* TMW Media Group.

*Atomic Theory of Matter: Part 1.* TMW Media Group.

*Atomic Theory of Matter: Part 2.* TMW Media Group.

*Atomic Theory of Matter: Part 3.* TMW Media Group.

*Atoms and Elements.* TMW Media Group.