Chapter 2: The Nature of Molecules

Chapter Synopsis

A basic understanding of chemistry is necessary to the study of biology because the two are inexorably intertwined. Living organisms are chemical machines composed of molecules that continually undergo chemical reactions to become new molecules.

Atoms are composed of protons, neutrons, and electrons. Each subatomic particle has its effect on the chemical identity and interactivity of each element with all other elements. Formation of molecules from elements depends primarily on the tendency of electrons to occur in pairs, balance positive and negative charges, and fill the outermost shell. Chemical bonds result from trading or sharing electrons; shared bonds are stronger because they require the continued close proximity of atoms to one another.

Water, a simple but elegant molecule, predominates in living organisms and is unique in the life-giving characteristics stemming from its polar nature. Water clings to other polar molecules (adhesion), as well as itself (cohesion), by forming transient hydrogen bonds. These bonds absorb thermal energy, consequently the presence of water has a moderating effect on temperature changes. It is also a powerful solvent for other polar molecules and excludes nonpolar molecules, enabling the formation of biological membranes. Water spontaneously dissociates into H^+ and OH^- ions and is the basis of most buffering systems.

Chapter Objectives

- Know the structure of an atom.
- Differentiate between atomic number and atomic mass.
- Understand the importance of and the differences between ions and isotopes of an element.
- Know what determines the chemical identity of a given element.
- Understand how molecules undergo reduction and oxidation reactions.
- Understand what controls the combination of elements into molecules.
- Know the characteristics of the various bonds found in living organisms.
- Explain why water is the essence of life as we know it.
- Describe pH in biological and mathematical terms. Know some values for household solutions.

Key Terms

acid
adhesion
anion
atom
atomic mass
atomic number
base
buffer
cation
chemical bond
chemical reaction
cohesion
compound
covalent bond
double bond
electron
electronegativity
element
energy level
half-life
heat of vaporization
hydration shell
hydrogen bond
hydrophilic
hydrophobic
hydrophobic exclusion
hydroxide ion
ion
ionic bond
ionic compound
ionization
inert
isotope
matter
molar concentration
mole
molecular formula
molecule
neutron
octet rule of eight
oxidation
pH scale

polar molecule	reactant reduction	surface tension
product	single bond	triple bond
proton	specific heat	valence electron
radioactive isotope	structural formula	water

CHAPTER OUTLINE

2.0 Introduction

I. ENORMOUS EXPLOSION MARKED THE BEGINNING OF THE UNIVERSE

A. Beginning of the Universe Began the Process of Evolution

B. Evolution of Life and of Molecules Utilized Same Processes fig 2.1

2.1 Atoms are nature's building materials

I. ATOMS

A. Universe Composed of Matter
1. All matter made of atoms
2. Difficult to study due to size

B. The Structure of Atoms
1. Composed of smaller subatomic particles fig 2.2
   a. Electrons ($^{-}$charge) in circular orbits around nucleus
      1) Same number as protons to balance charge
      2) Dictates chemical activity
   b. Protons ($^{+}$ charge) and neutrons (0 charge) in central nucleus
2. Atomic number = number of protons
   a. Neutrons and protons have same mass
   b. Only protons have electrical charge

C. Atomic Mass
1. Mass versus weight
   a. Mass is the amount of a substance
   b. Weight is the force of gravity exerted on it
   c. Atomic mass = mass of protons + mass of neutrons
2. Mass measured in daltons
   a. Proton or neutron is roughly 1 dalton
   b. Electron is 1/1840 dalton, practically mass-less

D. Isotopes
1. All atoms of an element have the same atomic number (proton number)
2. An element cannot be broken into other substances by chemical means
3. Isotopes of an element have:
   a. Same number of protons, different number of neutrons
   b. Same number of electrons, thus same chemical properties
4. Example: Carbon-12 (^{12}C) versus carbon-13 and carbon-14 fig 2.3
5. Unstable forms, like carbon-14, decay
   a. Emit radioactive energy
   b. Half-life = time for half of a sample's atoms to decay
   c. Potential harmful side effects, exposure must be limited

E. Electrons
   1. Electrically neutral atom has same number of electrons and protons
   2. Electron orbit maintained by electrical attraction
   3. In ions the number of electrons and protons are different
      a. Element that possesses a net electrical charge
      b. Positive charge if electron lost, a cation
      c. Negative charge if electron gained, an anion

II. ELECTRONS DETERMINE THE CHEMICAL BEHAVIOR OF ATOMS

A. Arrangement Determines Chemical Properties of Element
   1. Orbital describes probable, not actual, location
   2. Shapes differ — fig 2.4
      a. Inner *s* orbitals are spherical
      b. More distant *p* orbitals are dumbbell-shaped
      c. Maximum number of two electrons per orbital
   3. Orbitals extremely far away from nucleus, atom mostly empty space
   4. Nuclei of different atoms rarely contact one another
   5. Electrons interact, determine chemical behavior

B. Energy Within the Atom
   1. Electrons (–) are attracted to protons (+)
   2. Energy required to keep electrons in orbit
   3. Electron energy of position is potential energy — fig 2.5
      a. Moving electron away from nucleus
         1) Requires energy
         2) Electron then has more potential energy
      b. Moving electron toward nucleus
         1) Releases energy
         2) Electron then has less potential energy
   4. Exchange of electrons between molecules — fig 2.6
      a. Oxidation is a loss of electrons
      b. Reduction is a gain of electrons
      c. Chemical energy stored in electrons by oxidation-reduction reactions
   5. Energy level schematics — fig 2.7
      a. Electrons represented as concentric rings called energy levels
      b. Electrons in outer most rings hold more energy
      c. Don't confuse energy levels (energy amount) and electron orbitals (location)

2.2 The atoms of living things are among the smallest

I. KINDS OF ATOMS

A. 92 Naturally Occurring Elements
   1. Have different number of protons, different arrangement of electrons
   2. Mendeleev discovered pattern of chemical properties — fig 2.8

B. The Periodic Table
   1. Eight groups of repeating chemical properties
   2. Based on interactions of valence electrons in outer shell
   3. Maximum of eight electrons in outer shell of elements important to life
      a. Elements at maximum are inert, not reactive
      b. Elements with one less than maximum are highly reactive
   4. Octet rule (rule of eight) states that atoms want their outer shell full

C. Distribution of Elements in Living Organisms tbl 2.1
1. Only eleven elements found in greater than trace amounts
2. Elements are generally light, atomic mass less than 21

2.3 Chemical bonds hold molecules together

I. IONIC BONDS FORM CRYSTALS

A. Molecule Is a Stable Group of Atoms
1. Compounds are molecules containing more than one kind of element
2. A chemical bond is the holding force
3. Atoms attracted by opposite electrical charges: Ionic bonds

B. A Closer Look at Table Salt
1. Atoms donate or receive electrons from other atoms fig 2.9
2. Example: Sodium chloride, common table salt
   a. Sodium atom, loses electron = Na^+
   b. Chlorine atom, accepts electron = Cl^-
3. Resulting atoms become charged ions, an ionic compound
4. Bond forms by attraction of ions of opposite charges
   a. Not between two individual atoms
   b. Between one ion and all oppositely charged ions in vicinity
   c. Dissociate into ions when placed in water

II. COVALENT BONDS BUILD STABLE MOLECULES

A. Covalent Bonds
1. Two atoms share one or more pairs of valence electrons
2. Example: Single bonded diatomic hydrogen (H_2) fig 2.10
   a. Hydrogen has unpaired electron and unfilled outer level
   b. Two atoms combine, each nucleus shares two electrons
3. Bond requires close proximity of atoms to one another
4. Bond is very stable
   a. Has no net charge
   b. Octet rule satisfied, has no free electrons

B. Covalent Bonds Can be Very Strong
1. Strength of bond depends on number of shared electrons
   a. Double bond shares two pairs of electrons, stronger than a single bond
   b. Triple bond strongest covalent bond
2. Structural formulas: H – H or O = O
3. Molecular formulas: H_2 or O_2

C. Molecules with Several Covalent Bonds
1. Atoms can share electrons with more than one other atom
2. Example: Carbon has six electrons, four in the outer level
   a. Must gain four electrons to satisfy octet rule
   b. Thus can form four chemical bonds

D. Chemical Reactions
1. Formation and breaking of chemical bonds
2. Involve shifting atoms without change in number or identity
   a. Reactants: Original, pre-reaction molecules
   b. Products: Molecules resulting from a reaction

3. Influenced by several factors
    a. Temperature: Heat increases rate
    b. Concentration: Reactant versus product have opposite effect
    c. Catalyst: Special substance increases rate

2.4 Water is the cradle of life

I. CHEMISTRY OF WATER

A. Unique Properties of Water Necessary for Living Organisms
1. Exists as liquid at temperature of earth's surface — fig 2.11
2. Provides a medium in which other molecules can interact
3. Composes two-thirds of most organisms

B. The Atomic Structure of Water
1. Simple atomic structure, H_2O — fig 2.12a
2. Forms chemical bonds much weaker than covalent bonds
    a. Property is derived from structure of water
    b. Responsible for organization of living chemistry

II. WATER ATOMS ACT LIKE TINY MAGNETS

A. Electronegativity Attracts Electrons of Water Molecules
1. Oxygen atom more electronegative than hydrogen atom
2. Electrons attracted more strongly to oxygen than to hydrogen

B. Charge Separation Results in Polar Nature
1. Has distinct ends, each with a partial charge
2. Most stable configuration is tetrahedron, bond angle 104.5° — fig 2.12b
    a. Partial (δ+) charges at apexes opposite hydrogens
    b. Partial (δ–) charge at oxygen
3. Polar molecule results from magnet-like poles
4. Polarity is crux of chemistry of water and life
5. Polar molecules interact with one another
    a. Opposite charges attract, form hydrogen bonds — fig 2.13
    b. Bonds are transient, cumulative effects important
    c. Hydrogen bonds affect physical properties of water — tbl 2.2

III. WATER CLINGS TO POLAR MOLECULES

A. Polarity of Water Attracts It to Other Polar Molecules
1. Cohesion is attraction of water to water
    a. Results in surface tension of water — fig 2.14
    b. Causes things to get wet in water
2. Adhesion is attraction of water to another molecule
    a. Results in capillary action, water rises in thin tube — fig 2.15
    b. Attraction is electrostatic
    c. Height inversely proportional to tube diameter

B. Water Stores Heat
1. Exhibits high specific heat
    a. Amount of heat to change temperature of a substance
    b. Associated with and proportional to polarity
    c. Thermal energy must first disrupt hydrogen bonds

  - d. Heats up slowly, retains heat longer than surroundings
2. High heat of vaporization
   - a. Amount of heat required to change water to vapor
   - b. Evaporation of water produces cooling effect
3. Forms ice with decrease in temperature
   - a. Crystal-like lattice of hydrogen bonds — fig 2.16
   - b. Less dense than liquid water

C. Water Is a Powerful Solvent
1. Has ability to form hydrogen bonds
2. Water molecules gather around charged molecules
3. Example: Table sugar (sucrose)
   - a. Water forms hydrogen bonds with OH^- groups of sucrose
   - b. Each sugar molecule surrounded by cloud of water molecules
   - c. Cloud is called the hydration shell
4. Hydration shells also form around ions — fig 2.17

D. Water Organizes Nonpolar Molecules
1. Water excludes nonpolar molecules
2. Preferentially forms maximum number of hydrogen bonds
3. Minimizes disruption of hydrogen bonding
   - a. Hydrophobic: Not soluble in water, nonpolar
   - b. Hydrophilic: Soluble in water, polar
4. Hydrophobic exclusion
   - a. Forces nonpolar molecules to associate together
   - b. Shapes molecules with nonpolar regions

IV. WATER IONIZES

A. Water Covalent Bonds May Break Spontaneously
1. Proton dissociates from molecule
   - a. Becomes positively charged ion (H^+)
   - b. Remainder of molecule is hydroxide ion, OH^-
2. Ionization: Process of spontaneous ion formation
3. Mole of a substance is its molecular mass
   - a. Corresponds to combined atomic mass of all molecules
   - b. Molar concentration of H^+ ions in water is 10^{-7} mole / liter

B. pH
1. pH scale quantifies H^+ concentration — fig 2.18
   - a. pH = negative log of H^+ ion concentration = $-\log[H^+]$
   - b. pH of 7 indicates neutrality H^+ ions = OH^- ions
2. Scale is logarithmic, change of one on scale is really tenfold

C. Acids
1. Substance that dissociates to increase concentration of H^+
2. Has low pH value, below 7
3. Stronger acids have more H^+ ions

D. Bases
1. Substance that combines with H^+ ions when dissolved in water
2. Lower concentration of H^+
3. Also called alkaline solutions, have pH value above 7

E. Buffers
  1. pH of body fluids is about 7
  2. Minimize changes in H^+ and OH^- concentration — fig 2.19
  3. Act as reservoirs for H^+
     a. Donate to solutions when concentration falls
     b. Take from solutions when concentration increases
  4. Example: Carbonic acid / bicarbonate in blood — fig 2.20

INSTRUCTIONAL STRATEGY

PRESENTATION ASSISTANCE:

This is the material that many prospective biologists abhor. After all, if they enjoyed this stuff they would be taking chemistry as an elective, not biology. Although most programs consider basic high school chemistry a prerequisite to introductory biology, fewer high schools offer such a course now than did ten years ago. As a result, part of the class will be bored if you get too basic and the other part of the class will be lost if you assume this chapter is a review. Try to find a happy medium. A short pretest on the material may help gauge the level of your students, and may surprise some who thought they knew the material.

Many students have a math phobia as well as a chemistry phobia and have a difficult time with anything that has equations, plus, minus, and equal signs. pH is a difficult concept partly because of the invention of calculators; logarithms are ancient history. Stress that each number on the pH scale is different from its nearest neighbor by a factor of ten, like the Richter scale for earthquakes and the decibel scale for sound. Oxidation / reduction reactions cause problems as well; remember that reduced compounds add electrons and oxidized compounds lose electrons. This is one time that being reduced results in a gain!

Energy levels are like being on a pogo stick, you are either up or down, not in between. Electrons can only change their energy in specific increments, by being up or down. You can't accumulate bouncing – do a bounce action a few times and not move, and then at one point go up three times as high.

The characteristics of water are intuitive when related to everyday events, tempering effects on weather, sweating, surface tension, and so forth. Use as many common examples as possible. Your students can measure the relative pH of various household solutions using tea – the normal unadulterated drinking variety. Tea becomes more yellow in color when lemon juice is added because the juice is acidic, not because the tea is diluted by a yellow liquid. Red cabbage is also an acid-base indicator, red when acid, blue when basic.

VISUAL RESOURCES:

Molecular models of some sort are quite helpful. Many aspects of chemistry just don't work on a two-dimensional surface. Use students and an object to illustrate the difference between ionic and covalent bonds. When the object is given by one student to another, the recipient can walk away, no strings attached. When the object is to be shared, analogous to the covalent bond, the two students must remain in fairly close proximity for such sharing to be practical.

In a small class setting, bring in samples of polar and nonpolar substances and mix them together. In a large class, use an overhead to project it to the entire class; this may take a little ingenuity as you will be working on a horizontal surface. Cohesion and adhesion can also be demonstrated in this manner. Petri dishes and food coloring may help. Diatec makes 35 mm deep well projection slides that are waterproof (available through Carolina or Wards Biologicals).

The following analogy has been quite helpful in differentiating ionic and covalent bonds. Mary is a well-prepared student who sits attentively in the front row during lecture. Normally she brings two cans of pop to lecture, orange and cola. Ann, a thirsty classmate, begs the cola from generous

Mary and sits in the back row. The bond between the two students is analogous to an ionic bond. The can of pop is donated from one student to another. The bond strength between Mary and Ann is not very strong as they can sit on opposite sides of the lecture hall and still each drink a pop. David also comes to class with two cans of pop, root beer and lemon-lime. He, though, is less generous and less decisive than Mary and wants to drink both flavors of pop during lecture. When his thirsty friend Ed arrives, David decides to share his pop rather than overtly giving one can away. Ed must, therefore, sit in the seat right next to David. This is analogous to a covalent bond. David and Ed must remain in close proximity to one another and the bond between them is quite strong, especially in comparison to the ionic bond between Mary and Ann.