Biological Science Canadian 2nd Edition Freeman Solutions Manual

Chapter 2 – Water and Carbon: The Chemical Basis of Life

Learning Objectives: *Students should be able to...*

- Explain how and why atoms interact to form molecules. Sketch examples of how electrons are shared in nonpolar covalent bonds, polar covalent bonds, and ionic bonds.
- List the unusual properties of water. Explain how these properties relate to the structure of water molecules, and how they make water important for life.
- Define energy, and describe the major forms it can take. Explain why chemical bonds can be considered a form of energy.
- Explain, both in mathematical terms and in plain English, how changes in entropy and potential energy determine whether or not a reaction is spontaneous. Understand under what conditions a nonspontaneous reaction can occur.
- Know why carbon is a key element for life. List the six major functional groups, their structural formulas, and their basic characteristics.

Lecture Outline

I. Atoms, Ions, and Molecules: The Building Blocks of Chemical Evolution

- 96% of every organism is composed of the elements C, H, O, and N.
- Structure affects function; the physical structure of C, H, O and N affects the molecules that they form.

A. <u>Basic atomic structure</u>
1. Parts of an atom
    a. Protons: large, in nucleus, positive charge
        (1) The number of protons gives an atom its chemical identity.
        (2) Number of protons = atomic number
    b. Neutrons: large, in nucleus, no charge
        (1) The number of neutrons does not affect the atom's chemical identity but affects its mass.
        (2) Number of protons + number of neutrons = mass number
        (3) Isotopes are forms of an element with different numbers of neutrons.
    c. Electrons: small, outside nucleus, negative charge
    d. Most of an atom's volume is empty space. **(Fig. 2.1)**
2. Electron orbitals and valence
    a. Electrons occupy orbitals in energy shells around the nucleus.
    b. Each energy shell contains a specific number of orbitals. An orbital can hold up to two electrons.
    c. Unpaired electrons are unstable and tend to form bonds.
    d. Number of unpaired electrons = atom's valence **(Fig. 2.3)**

B. <u>How does covalent bonding hold molecules together?</u>
1. A pair of electrons is shared between two atoms, with each electron

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attracted to the protons of both nuclei. **(Fig. 2.4)**

   a. ***Students should be able to draw an analogy between atoms sharing electrons in a covalent bond and children arguing over a toy.***

2. If the electrons are shared equally between the two atoms, then the covalent bond is nonpolar. Examples: H–H bonds and C–H bonds. **(Fig. 2.5a)**
3. If the electrons are pulled closer to one of the atom's nuclei, then the covalent bond is polar.
   a. The tendency of an atom to hold electrons tightly is its electronegativity. Example: oxygen is more electronegative than hydrogen. **(Fig. 2.5b)**
   b. Polar covalent bonds result in partial charges on certain parts of the molecule. Example: water. **(Fig. 2.5b)**
   c. ***Students should be able to extend the toy analogy to polar and nonpolar covalent bonds***.
4. Single bonds share one pair of electrons, double bonds share two pairs, and triple bonds share three pairs. **(Fig. 2.8)**

C. <u>Ionic bonding, ions, and the electron-sharing continuum</u>
   1. In an ionic bond, the electron is lost entirely from one atom and donated to the other atom.
   2. The donor atom carries a (+) charge (cation); the recipient atom carries a (-) charge (anion). Example: sodium chloride. **(Fig. 2.6)**
   3. Nonpolar covalent bonds and ionic bonds represent two extremes of an electron-sharing continuum. **(Fig. 2.7)**
   4. ***Students should be able to extend the toy analogy to ionic bonds.***

D. <u>The geometry of simple molecules</u>
   1. The orientation of the orbitals containing shared electrons determines the angle of the bond, which affects the overall shape of the molecule. **(Fig. 2.9)**
   2. A molecule's geometry—its shape—affects its function. Example: water.

E. <u>Representing molecules</u> **(Fig. 2.10)**
   1. Molecular formula shows types and numbers of atoms in a molecule.
   2. Structural formula shows bonds between atoms.
   3. Ball-and-stick model is a 3-D representation showing bond geometry and indicating the relative size of the atoms.
   4. Space-filling model is the most accurate 3-D spatial depiction of the spatial relationship between atoms.

F. <u>Basic concepts in chemical reactions</u>
   1. A chemical reaction is a rearrangement of the bonds between atoms and can result in the breakdown of molecules or the formation of new molecules.
   2. Quantifying molecules:
      a. Mole = 6.022×10^{23} molecules (Avogadro's number)
      b. Molarity = number of moles per liter of solution

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II. The Early Oceans and the Properties of Water

- Cells are over 75% water. **(Fig. 2.11)**
- Water is an excellent solvent; most of the chemical reactions on which life depends take place between substances that are dissolved in water.

A. Why is water such an efficient solvent?
   1. Polarity and shape
      a. H_2O is a polar molecule due to the high electronegativity of oxygen.
      b. H_2O is bent, which allows the partial negative charge on the oxygen to "stick out."
   2. Hydrogen bonds
      a. Due to water's polarity and shape, hydrogen bonds can link regions of partial negative and positive charge on adjacent H_2O molecules. **(Fig. 2.12b)**
      b. ***Students should be able to draw water as a (hypothetical) linear molecule and describe why hydrogen bonds between linear water molecules would be much weaker.***
      c. Hydrophilic molecules can form hydrogen bonds with water and will dissolve in water.
         (1) Polar molecules and ions are hydrophilic. **(Fig. 2.13a)**
      d. Hydrophobic molecules cannot form hydrogen bonds with water and will not dissolve in water.
         (1) Nonpolar molecules are hydrophobic. **(Fig. 2.13b)**

B. How does water's structure correlate with its properties?
   1. Cohesion, adhesion, and surface tension result from water hydrogen-bonding with itself or with other hydrophilic substances. **(Fig. 2.14)**
   2. Water expands as it forms a solid, so ice floats. The bottom layers of cold lakes and oceans tend to remain unfrozen, allowing life to survive. **(Fig. 2.15)**
   3. Water has a very high specific heat and heat of vapourization, giving it insulating capabilities and cooling properties. **(Table 2.1)**

C. Acid–base reactions involve a transfer of protons.
   1. Dissociation of water: $H_2O \leftrightarrows H^+ + OH^-$
   2. Acids donate protons during a chemical reaction; bases take up protons.
   3. pH is a measure of the concentration of hydrogen ions in solution.
      a. $pH = -\log [H^+]$
      b. A pH below 7 is acidic; a pH above 7 is alkaline; a pH of 7 is neutral. **(Fig. 2.16)**
      c. ***Students should be comfortable with simple logarithms (BioSkills 7 in Appendix B) and should be able to figure out, without using a calculator, the pH of a solution if given a simple concentration of hydrogen ions (e.g., What is the pH of a solution that has 0.001 mole of hydrogen ions per liter?).***
   4. Buffers (weak acids) protect cells against damaging changes in pH by taking up H^+ ions when they are in excess and releasing them when

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they are scarce.

III. Chemical Reactions, Chemical Evolution, and Chemical Energy

A. How do chemical reactions happen?

1. In a chemical reaction, reactants are converted into products.
2. Most reactions are reversible.
3. Chemical equilibrium:
   a. Rate of forward reaction equals rate of reverse reaction; equilibrium is dynamic but stable.
   b. Equilibrium can be disturbed by adding more reactant or product or by altering the temperature.
4. Reactions that absorb heat are endothermic; reactions that release heat are exothermic.

B. What is energy?

1. Energy is the capacity to do work or supply heat.
2. Examples of different forms of energy:
   a. Potential energy = stored energy
      (1) The potential energy in chemical bonds is called chemical energy. **(Fig. 2.17)**
   b. Kinetic energy = energy of motion
      (1) The kinetic energy of molecular motion is called temperature.
3. First law of thermodynamics: Energy cannot be created or destroyed, but can be transferred or transformed. **(Fig. 2.18)**

C. Chemical evolution: a model system

1. Spontaneous chemical reactions are those that proceed on their own without any added energy.
   a. Reactions tend to be spontaneous if either or both of these conditions are met:
      (1) The products have lower potential energy than the reactants (negative ΔH; exothermic reaction). **(Fig. 2.19)**
      (2) The product molecules are less ordered than the reactant molecules (positive ΔS; increase in entropy). **(Figs. 2.20 and 2.21)**
2. The Gibbs free-energy change, ΔG, summarizes the combined effects of changes in heat and disorder.
   a. $\Delta G = \Delta H - T\Delta S$
   b. If ΔG is negative, the reaction is exergonic and spontaneous.
   c. If ΔG is positive, the reaction is endergonic and not spontaneous.
   d. If $\Delta G = 0$, the reaction is at equilibrium.
3. The role of temperature and concentration in chemical reactions:
   a. A spontaneous reaction may not be fast.
   b. Higher concentrations of reactants and high temperatures tend to speed up chemical reactions. **(Fig. 2.22)**
4. ***Students should be able to explain why the same reaction can be nonspontaneous at low temperature but spontaneous at high temperature.***
5. ***Students should be able to explain why some exothermic***

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reactions are nonspontaneous.

D. How did chemical energy change during chemical evolution?
   1. Early Earth was probably bombarded with high-energy photons, which can break molecules apart to form highly reactive free radicals. **(Fig. 2.23)**
      a. Free radicals are thought to be responsible for some of the key reactions in early chemical evolution.
   2. Experiments modeling conditions on early Earth indicate that formaldehyde (H_2CO) and hydrogen cyanide (HCN) would have formed.
      a. These are key intermediates in forming larger organic molecules.
   3. Chemical energy and chemical evolution
      a. During chemical evolution, the energy in sunlight was converted to chemical energy (the potential energy stored in chemical bonds).
      b. This allowed for larger, more organized molecules to be formed from smaller, simpler ones.

IV. The Importance of carbon

- A carbon atom can form four bonds, which enables carbon atoms to be linked in a wide variety of molecular shapes. **(Fig. 2.24)**
- Molecules that contain carbon are called organic molecules.
- Canadian Research 2.1: The carbon-rich Tagish Lake meteorite that landed in northern British Columbia in 2000 brought organic molecules, including amino acids, from space to Earth.

A. Linking carbon atoms together
   1. The carbon atoms in an organic molecule furnish a skeleton that gives the molecule its overall shape.
   2. After organic molecules formed on the early Earth, heat could drive the formation of more complex organic molecules from simpler ones.
      a. Examples: formaldehyde, acetaldehyde, and simple sugars. **(Fig. 2.25)**

B. Functional groups
   1. Functional groups are molecules added to a carbon skeleton that impart a variety of chemical reactivities to carbon molecules. **(Table 2.3)**
   2. There are six major functional groups:
      a. Amino groups function as bases.
      b. Carboxyl groups function as acids.
      c. Carbonyl groups are reactive with one another and can form C–C bonds.
      d. Hydroxyl groups are highly soluble in water and also act as weak acids.
      e. Phosphate groups have two negative charges and can affect the shape of the molecule. Large amounts of energy can be released when the bonds between adjacent phosphate groups are broken.
      f. Sulfhydryl groups can link two molecules via disulphide (S – S) bonds.

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Chapter Vocabulary

chemical evolution
carbon
nitrogen
hydrogen
oxygen
atom
nucleus
proton
neutron
electron
element
atomic number
isotope
mass number
dalton (Da)
orbital
electron shell
valence shell
valence electron
valence
chemical bond
molecule
covalent bond
electronegativity
nonpolar covalent bond
polar covalent bond
polarity
ionic bond
ion
cation
anion
single bond
double bond
triple bond
molecular formula
structural formula
ball-and-stick model
space-filling model
chemical reaction
mole
molecular weight
molarity
solution
solvent
solute
hydrogen bond
hydrophilic
hydrophobic
cohesion
adhesion
meniscus
surface tension
specific heat
heat of vapourization
hydrogen ion
hydroxide ion
acid
base
acid–base reaction
acidity
alkalinity
pH
pH scale
buffer
homeostasis
reactant
product
chemical equilibrium
system
concentration
endothermic
exothermic
energy
potential energy
kinetic energy
thermal energy
temperature
heat
first law of thermodynamics
chemical energy
sound
energymechanical
energy
electronegativity
spontaneous chemical reactions
nonspontaneous chemical reactions
ΔH
entropy
ΔS
second law of thermodynamics
Gibbs free energy
ΔG
exergonic
endergonic
photon
ozone
free radical
organic molecule
inorganic molecule
formaldehyde
hydrogen cyanide
acetaldehyde
functional group
amino group
amine
carbonyl group
aldehyde
ketone
carboxyl group
carboxylic acid
hydroxyl group
alcohol
phosphate group
organic phosphate
sulfhydryl group
thiol
disulphide bond
meteorite

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Lecture Activities

Lecture demonstrations

Estimated duration of activity: A few minutes during lecture for each demonstration

Polarity of water: If you have a sink in your lecture theatre, you can demonstrate water's polarity by turning on the tap and running a gentle stream of water. Take a glass rod and rub it on a small piece of fur to give it a static electrical charge. Move the charged rod towards the running water. The stream of water will bend towards the glass rod as a result of the partial charges within each water molecule.

Surface tension: If you are careful, you can place a thin needle on a water surface without it sinking.

Adhesion: If you touch a paper towel to a puddle of water, the towel will draw in the water. If you place Saran Wrap so that it touches the puddle, it will not draw in the water.

Student-Led Concept Illustrations

Estimated duration of activity: A few minutes during lecture

Often, getting students involved is as simple as having them describe a phenomenon in their own words. The following questions can be presented to students, allowing them to explain these concepts in terms that their peers can understand.

Types of chemical bonds: After describing hydrogen, covalent, and ionic bonds, ask the students to come up with nonbiological analogies to these types of bonds. You can start their thinking by comparing the molecules in covalent bonds to partners in a three-legged race. Another chemical-bond analogy that students tend to identify with is different levels of interpersonal relationships: Hydrogen bonds are analogous to acquaintances, ionic bonds are analogous to dating relationships, and covalent bonds are analogous to marriages. Ask the students to explain why each type of bond can be represented by each type of interpersonal relationship.

The properties of water: After explaining water's unique properties, ask students to explain the following phenomena:

- A leaf can land on the water's surface, but a rock sinks.
- There is less variation between day and night temperatures at the beach than between day and night temperatures in the desert.
- Salt dissolves in water, but gasoline does not.
- All mammals that live in hot environments keep themselves cool using two mechanisms: sweating and/or panting.

Energy: After introducing different forms of energy, ask students what sort of energy transformation is occurring in the following common daily activities. (Once they have the idea, you may challenge them to come up with additional examples.)

- Rubbing hands together to get warm (mechanical energy to thermal

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energy)
- Hearing (sound energy to mechanical energy to chemical energy; sound waves cause our eardrums to vibrate, triggering an electrical signal in our nervous system)
- A plant growing (electromagnetic energy to chemical energy; photosynthesis)
- A person riding a bicycle (chemical energy to mechanical energy)

Practicing BioSkills: Using Logarithms (BioSkills 7 in Appendix B)
Estimated duration of activity: 15 minutes
Many introductory students are not familiar with the concepts of logarithms, scientific notation, and molarity. The pH scale is a good opportunity to review all three concepts.

First, model for students how to convert a concentration of hydrogen ions (e.g., 0.001 mole of hydrogen ions per liter) to scientific notation (1×10^{-3}) and then to pH (3). Ask students why there is a 3 in the first number (i.e., in the number 0.001, the 1 is three places away from the decimal point). Then write the formal pH equation:

$$pH = -\log[H^+]$$

It may be useful to "translate" this mathematical equation into English by telling students that *log* is a formal way of asking "What's the exponent?"

When students understand the basic concepts of logarithms and scientific notation, have pairs of students work on these questions:
- What is the pH of a solution that has 0.01 mole of hydrogen ions per liter?
- What is the pH of a solution that has 0.0000001 mole of hydrogen ions per liter?
- A solution has a pH of 4. What is the molarity of H^+ ions?
- As the pH gets *higher,* does the concentration of hydrogen ions get lower or higher?
- A solution changes from pH 1 to pH 2. How much did the hydrogen ion concentration change?

Allow 5–10 minutes and then poll the class about their results, review the correct answers, and correct any misconceptions. Remind students that they can consult BioSkills 7 for further practice with logarithms.

Concept-Processing Pairs
Estimated duration of activity: 10–20 minutes (depending on the level of student understanding and the number of items used)
This activity has two parts. Students are divided into pairs. One member of each pair is designated A, and the other is B.
Part 1, Paired review: Give students 3–5 minutes to clarify with each other the basic concepts of the chapter, such as:
- Types of chemical bonds
- Polarity
- Properties of water
- Solubility

Part 2, Elaborative questions: Have students take turns asking each other questions that test understanding. Designate who will attempt to answer the first question (i.e., pick either student A or student B to start; this saves time for the student pairs with too much inertia to decide for themselves who will start). The number of questions asked depends on the length of time allotted for the activity. Make sure adequate time is provided for each question. Students will need time to process their answers, hear your optimal answer, and then discuss their understanding. The purpose of the exercise is to help students recognize and repair misconceptions and holes in their understanding.

A typical agenda for this activity:

1–2 minutes: Student A explains the answer to the question (as best as s/he can).
1–2 minutes: Student B assesses and amends the answer (as best as s/he can).
1 minute: Instructor shares the optimal answer with the class.
1–2 minutes: Student pairs discuss their answer relative to the answer.
1–3 minutes: Instructor entertains questions from the class.
Repeat the process with another question.

Sample questions:

- A molecule of octane contains about six times the mass of a molecule of water. Nevertheless, if one pours liquid octane onto liquid water, the octane floats on top of the water. How can this be?
- A salt crystal dropped into a beaker of water becomes smaller and eventually seems to disappear. However, the same salt crystal remains intact at the bottom of a beaker of octane. Explain.
- A beaker of water is allowed to sit underneath another larger beaker placed upside down over it (thereby trapping air above the water). After a long period of time, a few air molecules are found in the water and a few water molecules are present in the air above the beaker. However, the air and the water mostly remain separated. Why?
- (follow-up to the preceding question) Air contains mostly nitrogen (N_2) and oxygen (O_2). Explain how the structures of these two gases contribute to the observation in the preceding question.
- One liter of water is mixed with 0.2 liter of table salt. After the salt is dissolved, the new volume is slightly greater than 1 liter but much less than 1.2 liters. Why?

Discussion Idea: Calories and Body Temperature

Students like to relate what they are learning to the processes that occur in their own bodies. While discussing Gibbs free energy and the laws of thermodynamics, you can relate these concepts to nutrition and body temperature. Begin by asking students these two questions:

- Why do you heat up when you exercise?
- Why is exercising said to “burn calories”?

Give students two hints: (1) In the field of nutrition, the calorie is a unit of energy; that is, if one serving of a food contains 50 “calories,” that food will release 50

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kilocalories of energy when broken down by cells. (2) In any spontaneous reaction, at least some of the energy is converted into heat and is “lost”. This lost energy cannot be captured in chemical bonds. With those two hints, students should be able to reach the following conclusions:

- The oxidation of food releases energy that is captured by cells in the bonds of ATP; that is, chemical energy in food molecules is converted into chemical energy in ATP molecules. Some energy is lost as heat.
- When you exercise, the energy in the bonds of ATP is used to fuel muscle contraction. Again, some of this energy is lost as heat.
- The more molecules of ATP a muscle uses, the more energy is lost as heat. Muscles rapidly heat up. They also rapidly “run out” of ATP molecules and soon need more food molecules to produce more molecules of ATP. (Students may be interested to learn that at maximum contraction, a muscle will run out of ATP molecules in just a few seconds. This is why athletes can sustain a maximum contraction—such as in Olympic weight lifting—for only a few seconds.)

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