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

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

- Describe how and why atoms interact to form molecules. Sketch examples of how electron pairs are shared in nonpolar covalent bonds, polar covalent bonds, and ionic bonds.
- List the unique properties of water. Explain how these properties relate to the structure of water molecules.
- Explain how the structure of water explains its biologically important properties.
- Define energy and describe some of the major forms that energy can take. Explain why chemical bonds can be considered a form of potential energy.
- Explain in simple terms how changes in entropy and potential energy determine whether or not a reaction is spontaneous.
- Explain why carbon is a key element for life on Earth. 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 all living matter 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; shared electrons are attracted to the protons of both nuclei. Sharing of electrons can fill the

outer valence shell of atoms. **(Fig. 2.4)**

2. If the electrons are shared equally between the two atoms, then the covalent bond is nonpolar. Examples: H–H bonds **(Fig. 2.5a)** and C–H bonds.
3. If the two atoms do not share the pair of electrons equally, 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, such as water. **(Fig. 2.5b)**
4. Single bonds share one pair of electrons, double bonds share two pairs, and triple bonds share three pairs. **(Fig. 2.8).**

C. Ionic bonding, ions, and the electron-sharing continuum
1. In an ionic bond, the electron pair is transferred from one atom to the other atom.
2. The donor atom carries a (+) charge (cation), while 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 draw arrows between the atoms in each molecule shown in*** *Figure 2.8* ***to indicate the relative position of the shared electrons. If the electrons are equally shared, then students should draw a double-headed arrow.***

D. The geometry of simple molecules
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. Representing molecules **(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 relative sizes of atoms and the spatial relationship between atoms.

II. Properties of Water and the Early Oceans

- Cells are over 75% water.
- 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. H_2O is a polar molecule due to the high electronegativity of oxygen.
   a. H_2O is bent, which allows the partial negative charge on the oxygen to "stick out." **(Fig. 2.11)**

b. Due to water's polarity and shape, hydrogen bonds can link regions of partial negative and positive charge on adjacent H_2O molecules.
c. ***Students should be able to (a) draw a fictional version of** Figure 2.11a **that shows water as a linear (not bent) molecule with partial charges on the oxygen and hydrogen atoms and (b) explain why electrostatic attractions between such water molecules would be much weaker as a result.***
d. Hydrophilic molecules can form hydrogen bonds with water and will dissolve in water.
   (1) Polar molecules and ions are hydrophilic. **(Fig. 2.12)**
e. Hydrophobic molecules cannot form hydrogen bonds with water and will not dissolve in water.
   (1) Nonpolar molecules are hydrophobic. **(Fig. 2.13)**

B. <u>What properties are correlated with water's structure?</u>
1. Cohesion, adhesion, and surface tension result from water's ability to hydrogen-bond with other water molecules and 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)**

C. <u>The role of water in acid–base chemical reactions</u>
1. Dissociation of water: $2H_2O \leftrightarrows H_3O^+ + OH^-$
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 able to explain that a change of pH from 7 to 5 represents a solution that has 100-fold more protons and is 100 times more acidic.***
   d. ***Students should be able to calculate the concentration of protons in a solution that has a pH of 8.5.***
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 they are scarce.

III. Chemical Reactions, Energy, and Chemical Evolution

- The theory of chemical evolution suggests that simple molecules present on the early Earth participated in chemical reactions that eventually produced larger, more complex organic molecules.
- These reactions may have occurred in the atmosphere, which was dominated by gases ejected from volcanoes, or in deep-sea hydrothermal vents.

A. <u>How do chemical reactions happen?</u>
1. In a chemical reaction, reactants are converted into products.
2. Most reactions are reversible.

3. In a chemical equilibrium, the rate of the forward reaction equals the rate of the reverse reaction. The equilibrium is dynamic but stable.
   a. 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.
   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.
2. First law of thermodynamics: Energy cannot be created or destroyed, but can be transferred or transformed.

C. What makes a chemical reaction spontaneous?
1. Spontaneous chemical reactions are those that proceed on their own without added energy.
2. Reactions tend to be spontaneous if the products are less ordered than the reactants; example is explosion of nitroglycerin into CO_2, N_2, O_2, and H_2O with the release of heat.
   a. Amount of disorder in a system is called entropy.
   b. Entropy increases when products of reaction are less ordered than reactants.
   c. Second law of thermodynamics states that entropy increases in a closed system.
3. Reactions tend to be spontaneous if the products have a lower potential energy than reactants.
   a. If electrons are held more tightly in products than in the reactants, then the product has lower potential energy.
   b. Equal sharing of electrons in molecules such as H_2 or O_2 results in higher potential energy than in H_2O, where the electronegative oxygen holds the electrons more tightly than in O_2.
   c. In the reaction of $2H_2 + O_2 \rightarrow 2H_2O$, entropy decreases, but the potential energy of the products is much lower than the reactants, and a large release of potential energy as heat results in the reaction occurring spontaneously. **(Fig. 2.18)**
   d. Spontaneous processes result in lower potential energy, increased disorder, or both. **(Fig. 2.19)**

IV. Model Systems for Investigating Chemical Evolution

- To probe the kinds of reactions that set chemical evolution in motion, researchers focus on small molecules that were present on early Earth.
- Prebiotic soup model—molecules were synthesized from gases in atmosphere or on meteorites. These molecules ended up in the oceans upon condensation. Additional reactions resulted in more

complex molecules.
- Surface metabolism model—dissolved gases contacted minerals near deep-sea ocean vents and resulted in reactions generating more complex molecules.

A. Early origin-of-life experiments
   1. Graduate student Stanley Miller designed prebiotic soup experiment by simulating early Earth's atmosphere and ocean. **(Fig. 2.20)**
      a. Within a week, experiment produced highly reactive hydrogen cyanide and formaldehyde plus amino acids, the building blocks of proteins.

B. Recent origin-of-life experiments
   1. Early Earth was bombarded with high-energy photons, which can break up molecules to form free radicals. **(Fig. 2.21)**
   2. Free radicals are thought to have been responsible for some of the key reactions in early chemical evolution. They contain unpaired electrons in their outermost shells, making them highly reactive.
   3. Experiments modelling conditions on early Earth indicate that formaldehyde (H_2CO) and hydrogen cyanide (HCN) would have formed. Thus, key intermediates in forming larger organic molecules would have rained down into oceans. **(Fig. 2.22a)**
   4. Surface metabolism model accounts for the increased localized concentration of formaldehyde and hydrogen cyanide in Earth's early oceans.
      a. Reactants recruited to defined space—reactive minerals present on walls of deep-sea vent chimneys **(Fig. 2.22b)**
      b. Minerals play a critical role in rate of reactions as catalysts.
      c. Acetic acid can form under the conditions present at hydrothermal vents. Acetic acid is important in the synthesis of acetyl CoA (important throughout the tree of life).
      d. Evidence suggests that minerals from thermal vents were the original source of catalysts used in modern reactions.
   5. Canadian Research 2.1: Searching for life in extreme environments
      a. Scientists are searching for clues to ancient life in extreme environments on Earth.

V. The Importance of Organic Molecules

- Molecules that contain carbon are called organic molecules.
- A carbon atom can form four bonds, which enables carbon atoms to be linked in a wide variety of molecular shapes. **(Fig. 2.23)**
- Canadian Research 2.2: The carbon-rich Tagish Lake meteorite
   - A meteorite that landed in northern British Columbia in 2000 brought organic molecules from space to Earth.

A. Linking carbon atoms together
   1. The carbon atoms in an organic molecule form a skeleton that gives the molecule its overall shape.

B. Functional groups
   1. Functional groups are molecules added to a carbon skeleton that

impart a variety of chemical reactivities to carbon molecules. **(Table 2.1)**

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.

Chapter Vocabulary

carbon
nitrogen
hydrogen
oxygen
atom
nucleus
proton
neutron
electron
element
atomic number
atomic weight
radioactive isotope
mass number
dalton (Da)
orbital
electron shell
valence shell
valence electron
chemical bond
molecule
compound
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
heat of vapourization
hydrogen ion (H^+)
hydroxide ion (OH^-)
acid
base
acid–base reaction
acidity
alkalinity
pH
pH scale
buffer
homeostasis
reactant
product
chemical equilibrium
system
concentration
endothermic
exothermic
energy
potential energy
chemical energy
kinetic energy
thermal energy
temperature
heat
first law of thermodynamics
electronegativity
catalyst
spontaneous chemical reaction
nonspontaneous chemical reaction

entropy
second law of thermodynamics
photon
ozone
free radical
organic molecule
inorganic molecule
chemical evolution
prebiotic soup model
surface metabolism model
meteorite
formaldehyde
hydrogen cyanide
acetaldehyde
functional group
amino group
amine
carbonyl group
aldehyde
ketone
carboxyl group
carboxylic acid
hydroxyl group
alcohol
phosphate group
sulfhydryl group
thiol
disulphide bond

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: Rest a sheet of lens tissue on a water surface and then place a thin needle on the lens tissue. Sink the lens tissue to leave the needle supported by surface tension on the water surface.
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
Getting students involved may be 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 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.
- Mammals that live in hot environments can keep themselves cool by 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 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)

Using Logarithms

Estimated duration of activity: 15 minutes

Many introductory students will benefit from reviewing the concepts of logarithms, scientific notation, and molarity.

First, model for students how to convert a concentration of hydrogen ions (e.g., 0.001 mole of hydrogen ions per litre) 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 litre?
- What is the pH of a solution that has 0.0000001 mole of hydrogen ions per litre?
- 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.

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 as student pairs do not have 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 their misconceptions.

Typical agenda for this activity:

1–2 minutes: Student A explains the answer to the question to his/her best ability.

1–2 minutes: Student B amends the answer to his/her best ability.

1 minute: Instructor shares the optimal answer with the class.

1–2 minutes: Student pairs discuss their answer relative to the optimal 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 litre of water is mixed with 0.2 litre of table salt. After the salt is dissolved, the new volume is slightly greater than 1 litre but much less than 1.2 litres. 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 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 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 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.)

Biological Science

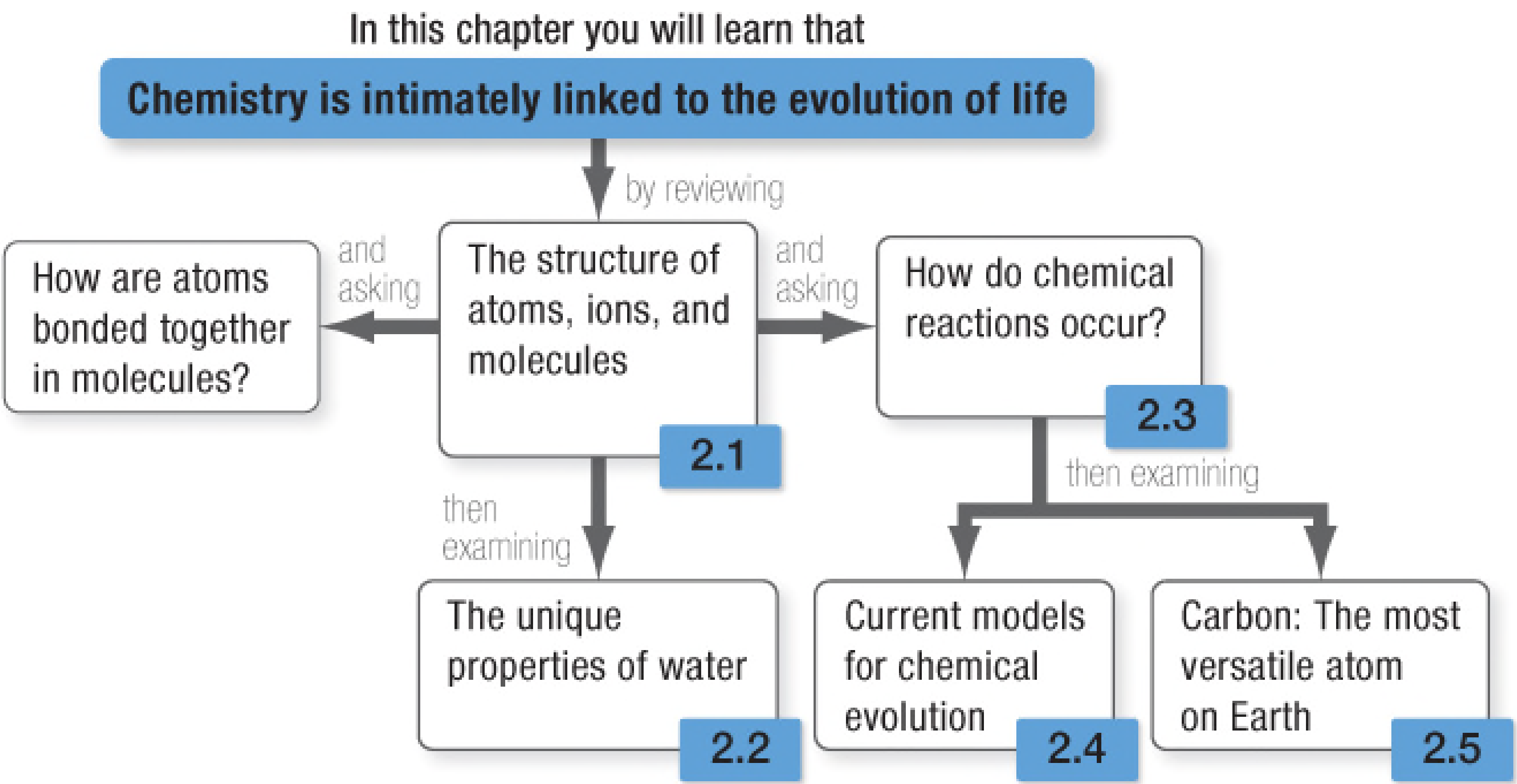
Third Canadian Edition

Chapter 2

Water and Carbon: The Chemical Basis of Life

Lectures by Cindy S. Malone, California State University Northridge, and Sharon Gillies, University of the Fraser Valley

Chapter 2 Opening Roadmap



Introduction to the Chemical Basis of Life (1 of 2)

- **Chemical evolution** is the leading explanation for the origin of life on Earth
  - Inputs of energy led to formation of increasingly complex carbon-containing molecules
  - Eventually led to a molecule that could replicate itself
  - Switch from chemical to biological evolution
- Evolution by natural selection took over
- A descendant of the original molecule became metabolically active and acquired a membrane
- Life had begun

Introduction to the Chemical Basis of Life

(2 of 2)

- Is the theory of chemical evolution plausible?
- What is the evidence?
- Let’s start with the atoms and molecules that would have combined to start chemical evolution

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

- Just four types of atoms make up 96% of matter in organisms—hydrogen, carbon, nitrogen, and oxygen
- To understand how simple substances evolve into complex structures in living cells, we must ask:
  1. What is the physical structure of hydrogen, carbon, nitrogen, and oxygen?
  2. What is the structure of the simple molecules—water, carbon dioxide, etc.—that served as the building blocks of chemical evolution?

Basic Atomic Structure (1 of 9)

- Atoms are composed of:
  - **Protons** – positively charged particles
  - **Neutrons** – neutral particles
  - **Electrons** – negatively charged particles
- Protons and neutrons are located in the **nucleus**.
- Electrons are found in **orbitals** surrounding the nucleus.

Figure 2.1 Parts of an Atom

(a) Diagrams of atoms

(b) Most of an atom's volume is empty space.

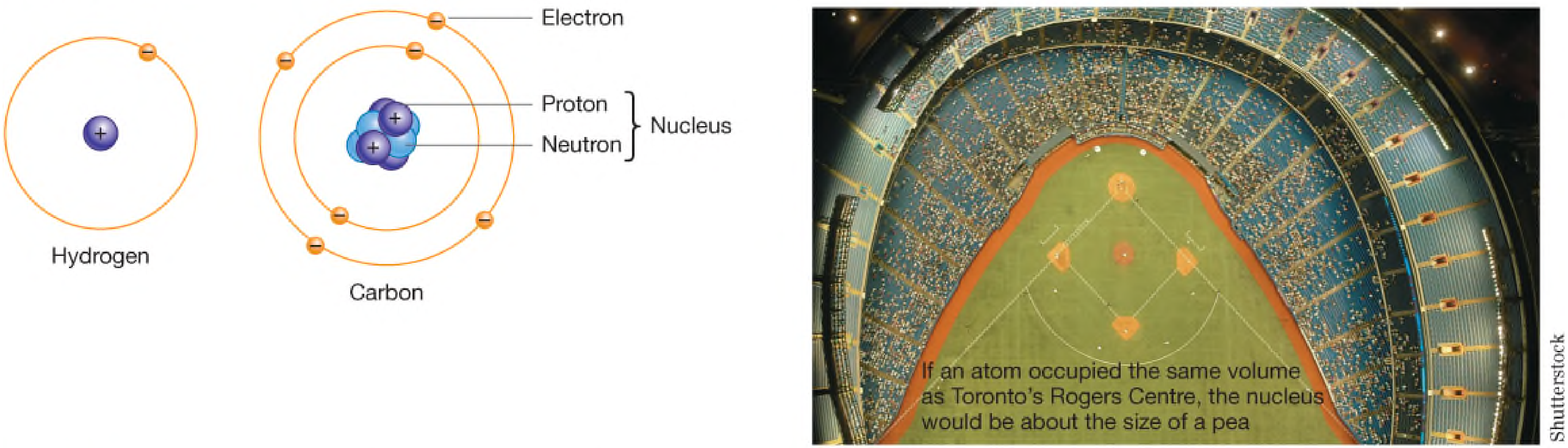


Figure 2.1 Parts of an Atom. A simplified model of an atom with its nucleus, made up of protons and neutrons—or a single proton in the case of hydrogen—surrounded by orbiting electrons. In reality, electrons are not evenly spaced, nor do they orbit the nucleus in concentric circles; their actual orbits are complex.

Basic Atomic Structure (2 of 9)

- The **atomic number**
  - The characteristic number of protons in the nucleus of any atom
  - Written as a subscript to the left of its symbol
- Atoms with the same atomic number
  - Have the same chemical properties
  - Belong to the same **element**

Figure 2.2 A Portion of the Periodic Table

Mass number (number of protons + neutrons)

Atomic number (number of protons)

$^{1}_{1}H$							$^{4}_{2}He$
$^{7}_{3}Li$	$^{9}_{4}Be$	$^{11}_{5}B$	$^{12}_{6}C$	$^{14}_{7}N$	$^{16}_{8}O$	$^{19}_{9}F$	$^{20}_{10}Ne$
$^{23}_{11}Na$	$^{24}_{12}Mg$	$^{27}_{13}Al$	$^{28}_{14}Si$	$^{31}_{15}P$	$^{32}_{16}S$	$^{35}_{17}Cl$	$^{40}_{18}Ar$

Figure 2.2 A Portion of the Periodic Table. Each element has a unique atomic number and is represented by a unique one- or two-letter symbol. The mass numbers given here are the most common for each element. (Appendix B provides a complete periodic table.)

Basic Atomic Structure (3 of 9)

- Protons have a +1 charge and electrons have a −1 charge
  - When an atom has an equal number of protons and electrons, the charges balance
  - The entire atom is electrically neutral

Basic Atomic Structure (4 of 9)

- The **mass number** is
  - The number of protons + neutrons in an atom
  - Written as a superscript to the left of its symbol
- Each proton and each neutron has a mass of one **Dalton (Da)**
- The mass of an electron is so small, it can be ignored
- Therefore, the mass of an atom is equal to its mass number

Basic Atomic Structure (5 of 9)

- **Isotopes** are
  - Forms of an element with different numbers of neutrons
  - Isotopes of an element have different masses
- Example: All carbon atoms have 6 protons
  - Carbon-12 has 6 neutrons; atomic mass 12 Da
  - Carbon-13 has 7 neutrons; atomic mass 13 Da
  - Carbon 14 has 8 neutrons; atomic mass 14 Da

Basic Atomic Structure (6 of 9)

- The **atomic weight** of an element
  - Average of all the masses of the naturally occurring isotopes based on their abundance
  - Example: The atomic number of carbon is 12.01, since carbon-12 is the most abundant isotope
- Most isotopes are stable, but some are unstable **radioactive isotopes** that decay over time

Basic Atomic Structure (7 of 9)

- Electrons move around atomic nuclei in specific regions called **orbitals**
  - Each orbital can hold up to two electrons
  - Orbitals are grouped into levels called **electron shells**
- Electron shells are numbered 1, 2, 3, and so on
  - Numbers indicate their relative distance from the nucleus
  - Smaller numbers are closer to the nucleus

Basic Atomic Structure (8 of 9)

- Each electron shell contains a specific number of orbitals
  - An electron shell comprising a single orbital can hold up to two electrons
  - A shell with four orbitals can contain up to eight electrons
- The electrons of an atom fill the innermost shells first and then fill the outer shells

Basic Atomic Structure (9 of 9)

- The outermost shell of an atom is the **valence shell**
- Electrons in this shell are **valence electrons**
- The number of unpaired valence electrons is called the **valence** of an atom
- Different atoms have different numbers of unpaired electrons

Figure 2.3 The Atomic Structure of the First 18 Elements

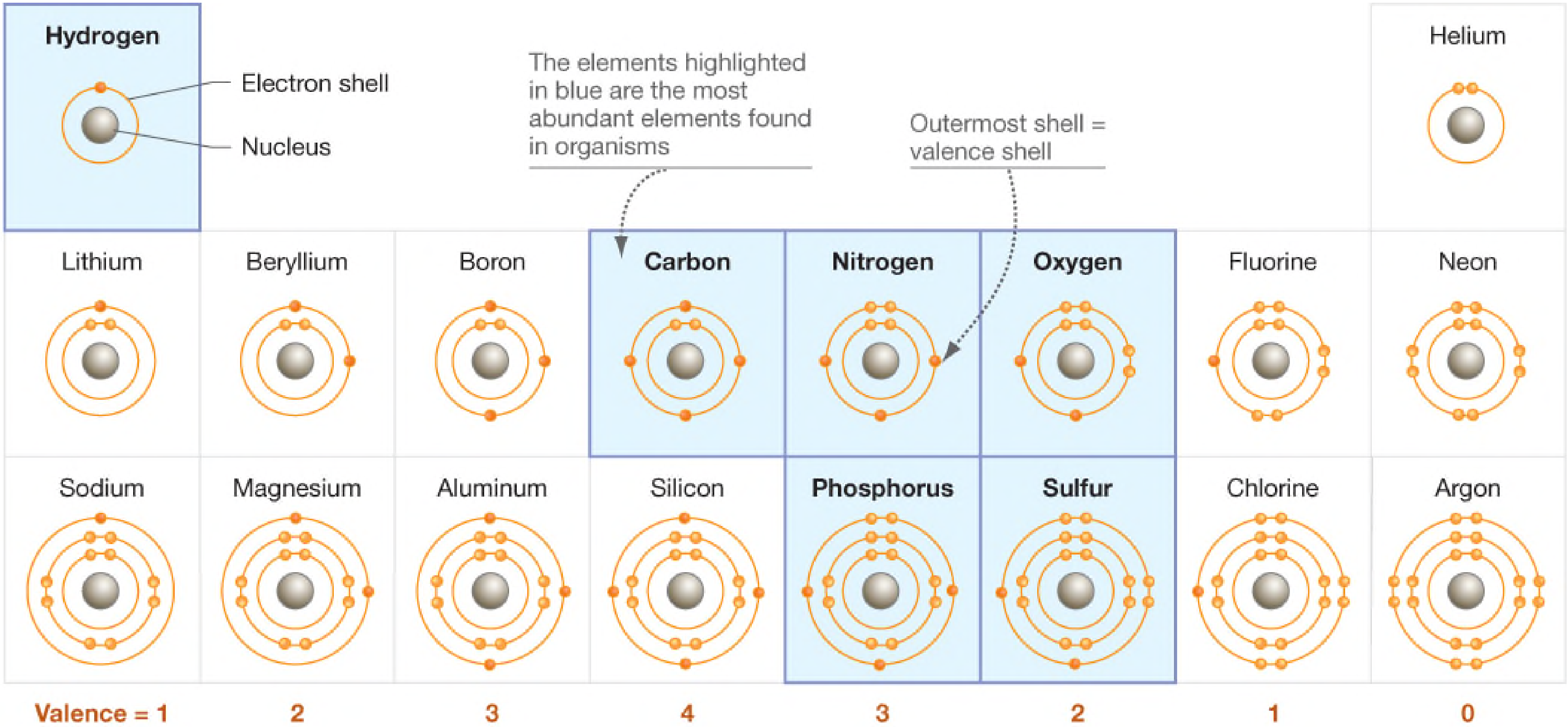


Figure 2.3 The Atomic Structure of the First 18 Elements. The most abundant elements in organisms are highlighted in blue.

How Does Covalent Bonding Hold Molecules Together? (1 of 2)

- Atoms are most stable when their valence shells are full
- Shells can be filled by formation of **chemical bonds**
  - Attractions that bind atoms together
  - **Covalent bonds** form when unpaired valence electrons are shared by two atoms
  - This effectively gives each atom a full outer shell

Figure 2.4 Covalent Bonds Result from Electron Sharing

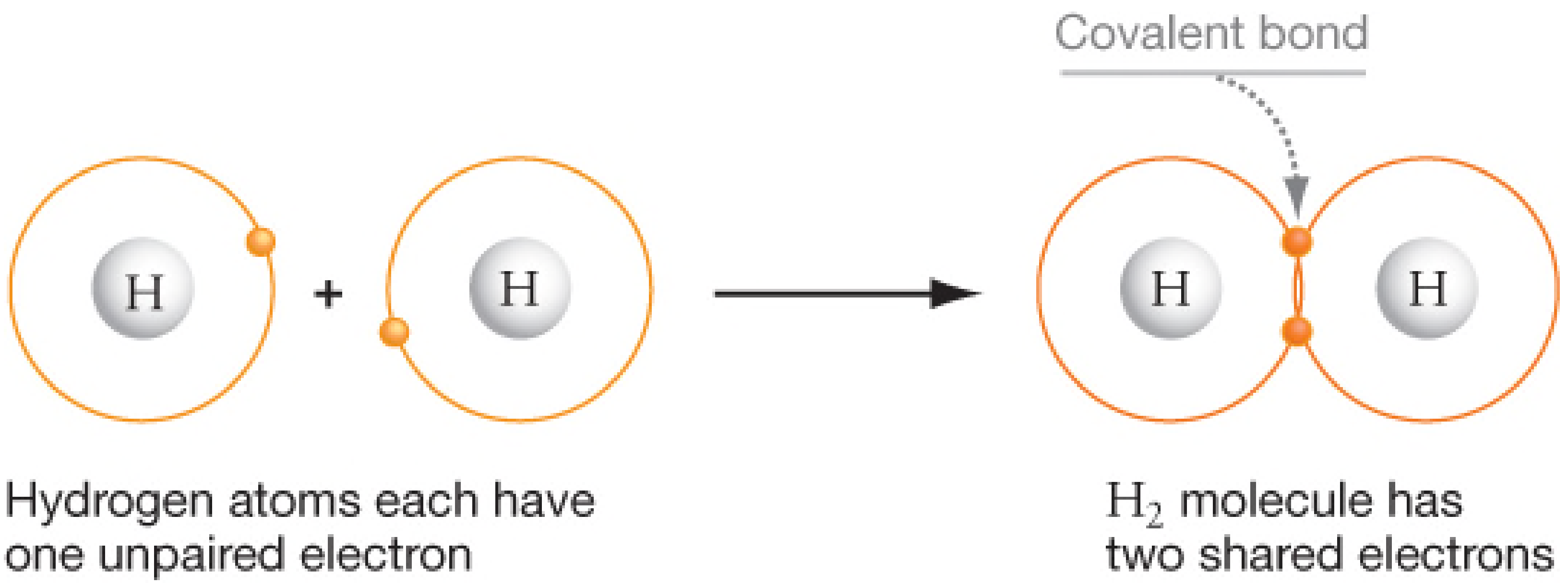


Figure 2.4 Covalent Bonds Result from Electron Sharing. When two hydrogen atoms form a covalent bond, their unpaired valence electrons are shared by each nucleus.

How Does Covalent Bonding Hold Molecules Together? (2 of 2)

- **Molecules** are substances held together by covalent bonds
- **Compounds** are molecules in which
  - atoms of different elements are held together

Nonpolar and Polar Bonds (1 of 2)

- Electrons are not always shared equally
- Atoms may have different **electronegativities**—the strength with which they pull electrons toward themselves
  - Determined by number of protons and the distance of the valence shell from the nucleus
  - In general, moving up and to the right on the periodic table = higher electronegativity
  - O > N > S,C,H,P

Nonpolar and Polar Bonds (2 of 2)

- Differences in electronegativity dictate how electrons are distributed in covalent bonds
  - **Nonpolar covalent bond**
    - Electrons are evenly shared between two atoms
    - The bond is symmetrical
    - Example: C–H bond
  - **Polar covalent bond**
    - Electrons are shared unevenly
    - Example: O–H bond

Figure 2.5 Electron Sharing and Bond Polarity

Figure 2.5 Electron Sharing and Bond Polarity. Electrons in a covalent bond can be **(a)** shared equally, resulting in nonpolar bonds, or **(b)** shared unequally, resulting in polar bonds. Delta symbols δ^+ and δ^- associated with polar covalent bonds refer to partial charges that arise owing to unequal electron sharing.

Polar Bonds Produce Partial Charges on Atoms

- An atom in a molecule with a high electronegativity
  - Holds electrons more tightly—has a partial negative charge (δ^-)
  - The other atom will have a partial positive charge (δ^+)

Ionic Bonding, Ions, and the Electron-Sharing Continuum (1 of 2)

- **Ionic bonds** result when electrons are transferred from one atom to another to give both atoms full valence shells
- An **ion** is an atom or molecule that carries a charge
  - **Cation—**an atom that loses an electron and becomes positively charged
  - **Anion—**an atom that gains an electron and becomes negatively charged
- Ionic bonds are the attraction between oppositely charged ions

Figure 2.6 Ion Formation and Ionic Bonding

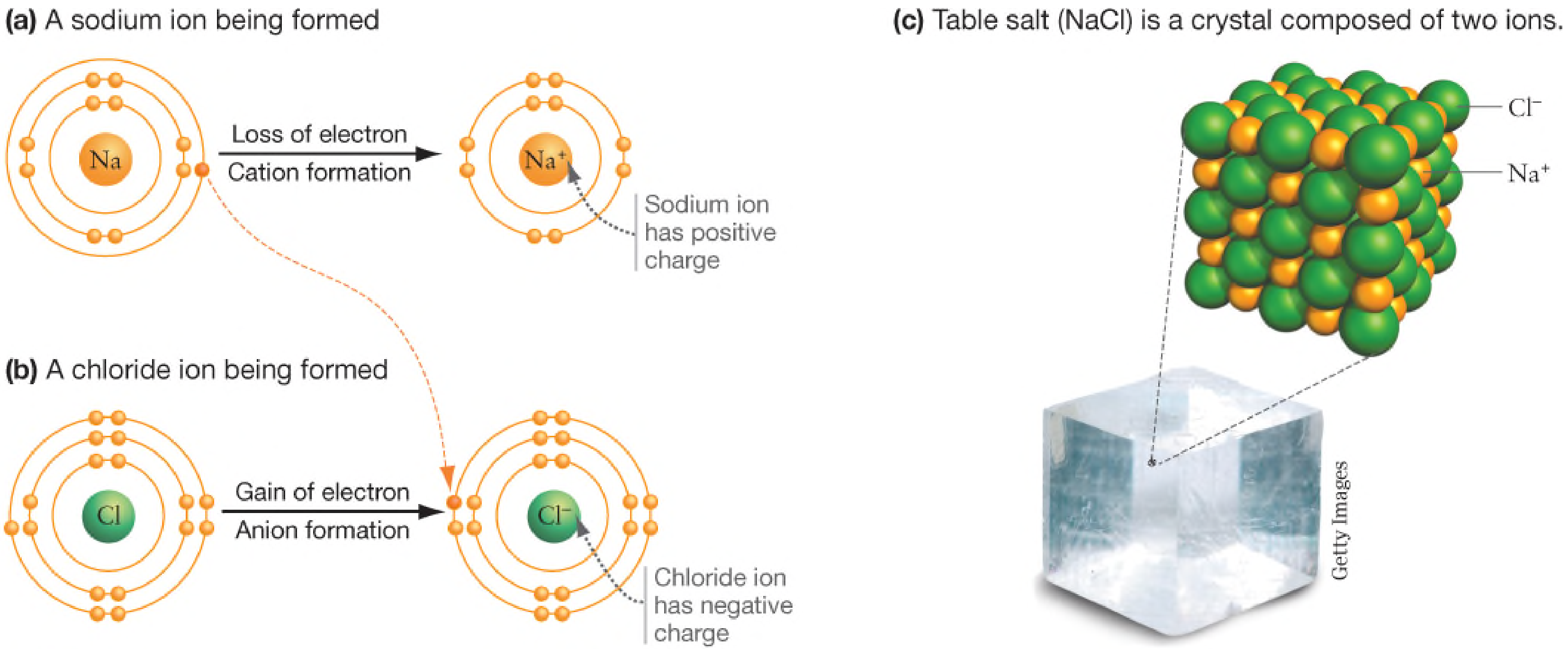


Figure 2.6 Ion Formation and Ionic Bonding. The sodium ion (Na^+) and the chloride ion (Cl^-) are stable because they have full valence shells. In table salt (NaCl), sodium and chloride ions pack into a crystal structure held together by electrical attraction between their positive and negative charges.

Ionic Bonding, Ions, and the Electron-Sharing Continuum (2 of 2)

- The degree to which electrons are shared in chemical bonds forms a continuum
  - From equal sharing in nonpolar covalent bonds
  - To unequal sharing in polar covalent bonds
  - To the complete transfer of electrons in ionic bonds

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Figure 2.7 The Electron-Sharing Continuum

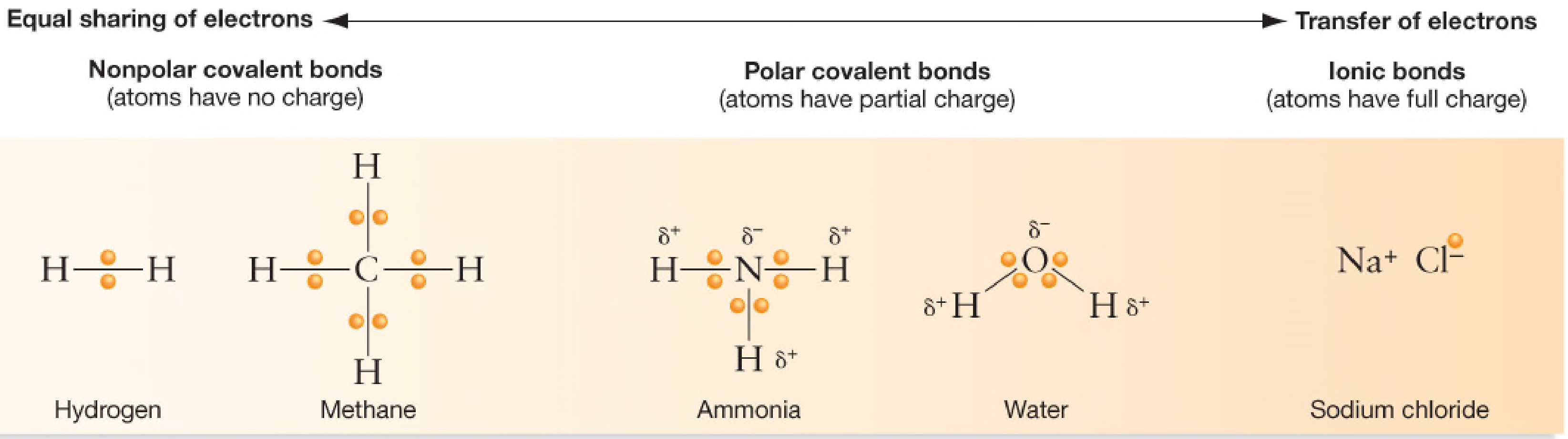


Figure 2.7 The Electron-Sharing Continuum. The degree of electron sharing in chemical bonds can be thought of as a continuum, from equal sharing in nonpolar covalent bonds to complete electron transfer in ionic bonds.

Some Simple Molecules Formed from C, H, N, and O

- The number of unpaired electrons determines the number of bonds an atom can make.
- Atoms with more than one unpaired electron
  - can form multiple single bonds
  - or double
  - or triple bonds.

Figure 2.8 Unpaired Electrons in the Valence Shell Participate in Covalent Bonds

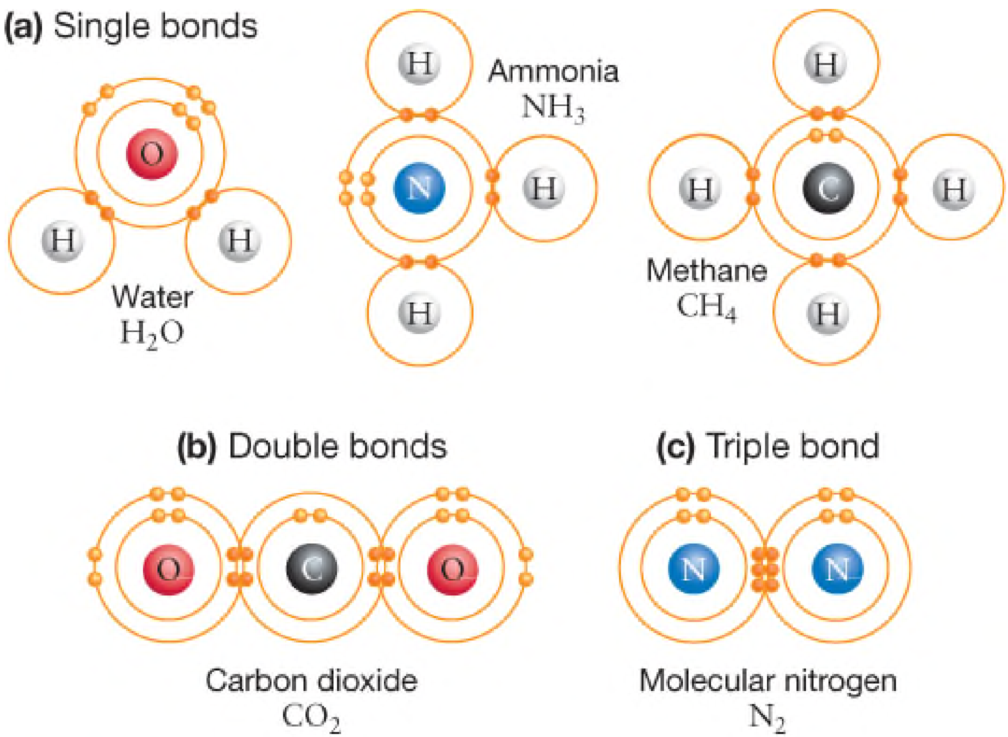


Figure 2.8 Unpaired Electrons in the Valence Shell Participate in Covalent Bonds. Covalent bonding is based on sharing of electrons in the outermost shell. Covalent bonds can be **(a)** single, **(b)** double, or **(c)** triple.

The Geometry of Simple Molecules

- A molecule's shape often dictates its behaviour
- The shape of a simple molecule is governed by the geometry of its bonds
  - Nitrogen (N_2) and carbon dioxide (CO_2) are linear
  - Methane (CH_4) is a tetrahedron since electrons repel each other and push as far apart as they can
  - Water (H_2O) is planar and bent because of the two unshared electron pairs

Figure 2.9 The Geometry of Methane and Water

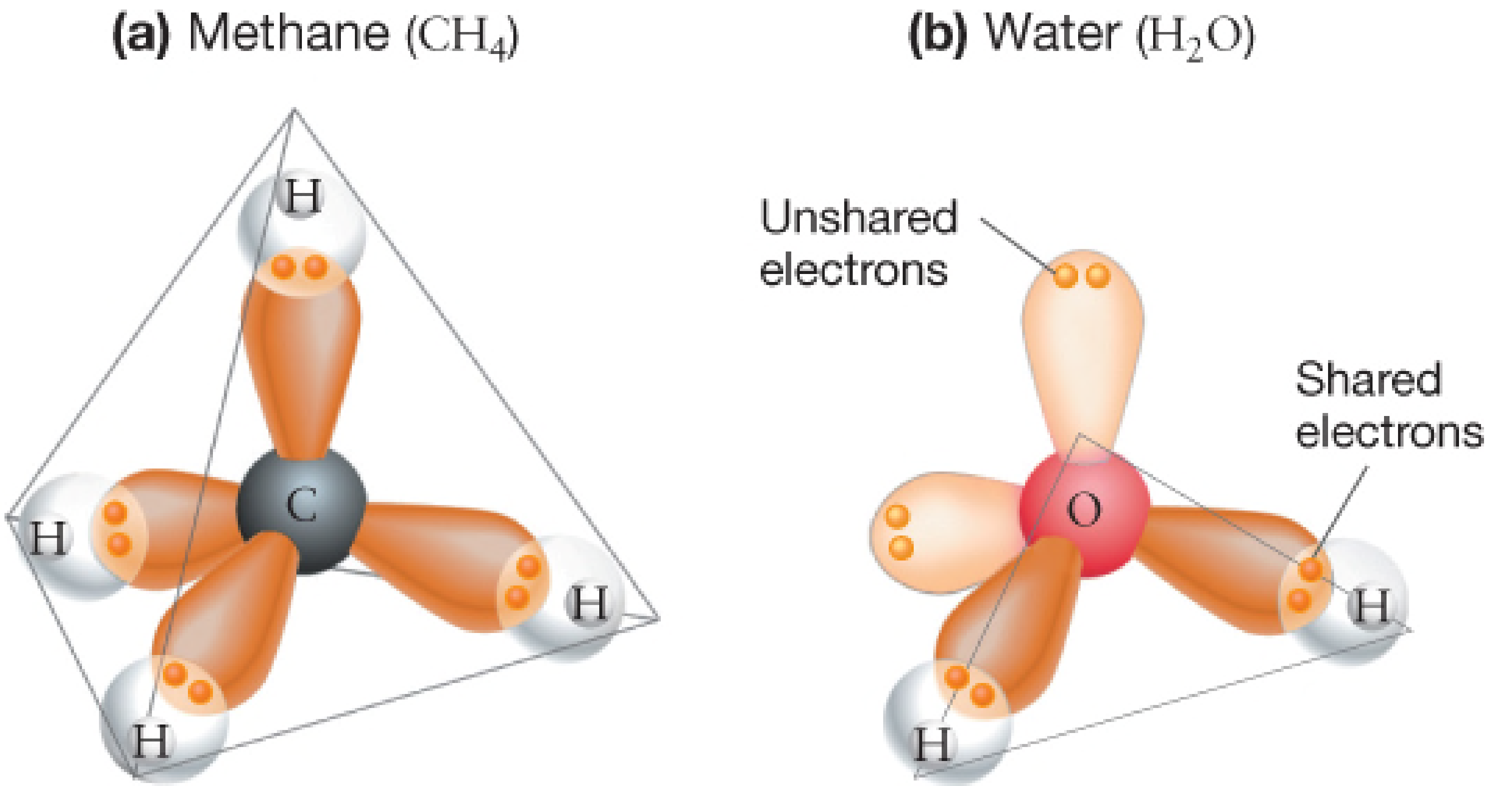


Figure 2.9 **The Geometry of Methane and Water.**

Representing Molecules

- Molecules can be represented in a variety of ways:
- **Molecular formulas** indicate the numbers and types of atoms in a molecule
  - Example H_2O, CH_4
- **Structural formulas** indicate which atoms are bonded together and whether the bonds are single, double, or triple bonds.
- **Ball-and-stick models** and **space-filling models** show 3D geometry.

Figure 2.10 Molecules Can Be Represented Several Ways

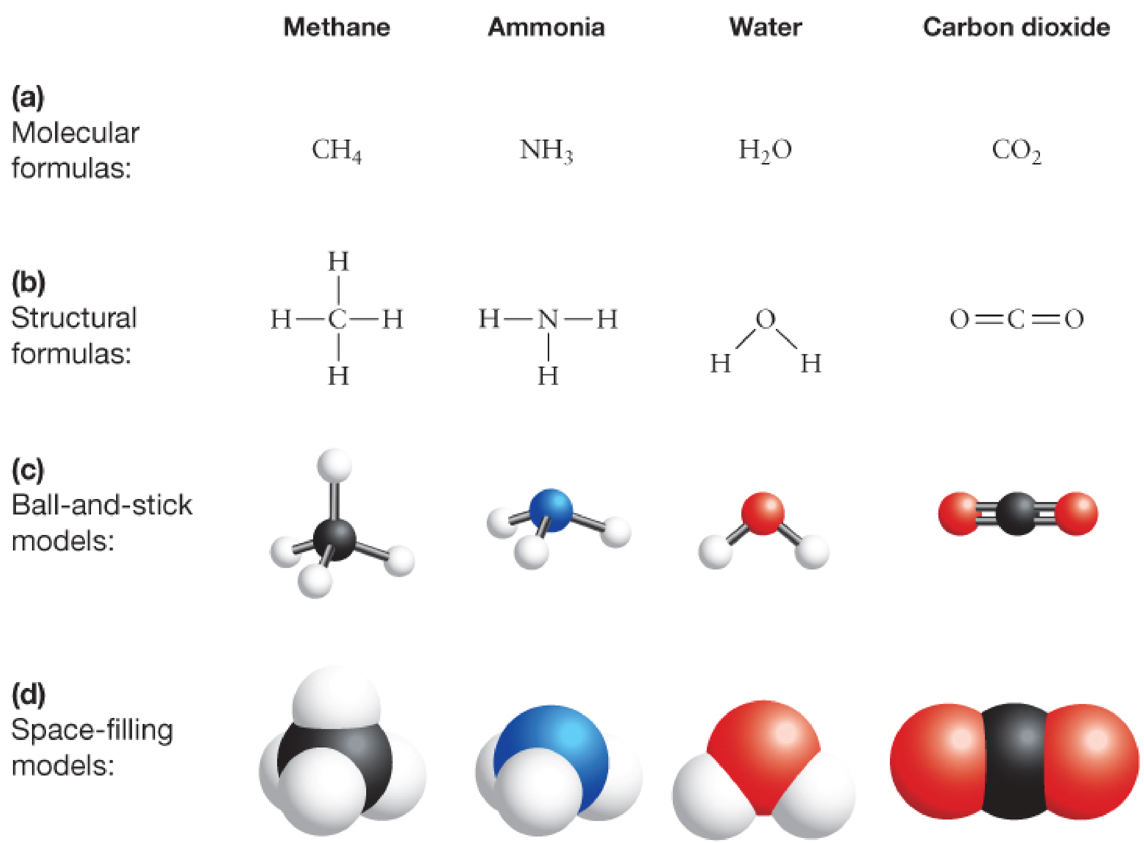


Figure 2.10 Molecules Can Be Represented Several Ways. Each method of representing a molecule has particular advantages.

Properties of Water and the Early Oceans

- Chemical evolution likely occurred in an aqueous, or water-based environment
- Life is based on water because water is an excellent **solvent**
  - A **solute** dissolved into a solvent makes a solution
  - Substances are more likely to react when they are dissolved in a solvent like water

Why Is Water Such an Efficient Solvent?

- Water is **polar**
  - The oxygen atoms have a partial negative charge
  - The hydrogen atoms have a partial positive charge
  - The charges are at opposite ends of a water molecule
- Water molecules interact with each other
  - The partial negative charges on oxygen
  - Attract the partial positive charges on hydrogen
  - These weak electrical attractions are called **hydrogen bonds**

Figure 2.11 Water Is Polar and Participates in Hydrogen Bonds

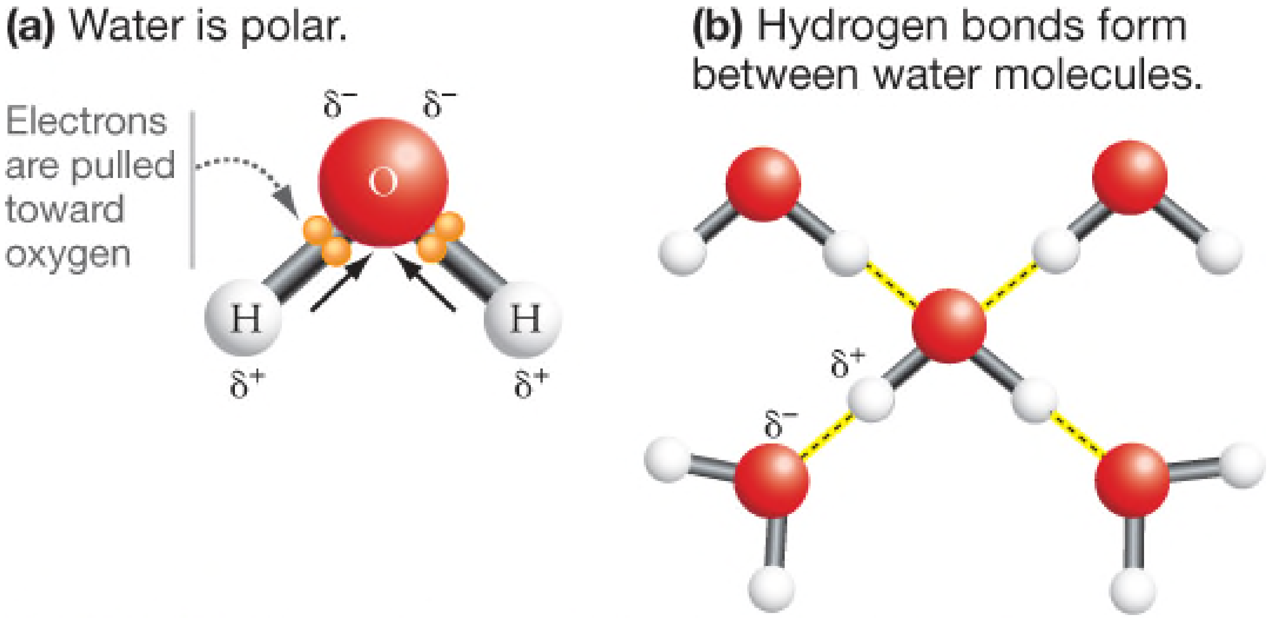


Figure 2.11 Water Is Polar and Participates in Hydrogen Bonds. **(a)** The polar covalent bonds in water give the oxygen a partial negative charge and each hydrogen atom a partial positive charge. **(b)** The partial charges on water molecules can form up to four hydrogen bonds. The oxygen can form two; each hydrogen can form one.

Why Is Water Such an Efficient Solvent?

- Hydrogen bonds can also form between a water molecule and any other polar molecule
- **Hydrophilic** (“water-loving”) atoms and molecules
  - Are ions and polar molecules that stay in solution
  - They interact with water’s partial charges
- Hydrogen bonding makes it possible for almost any charged or polar molecule to dissolve in water

Figure 2.12 Polar Molecules and Ionic Compounds Dissolve Readily in Water

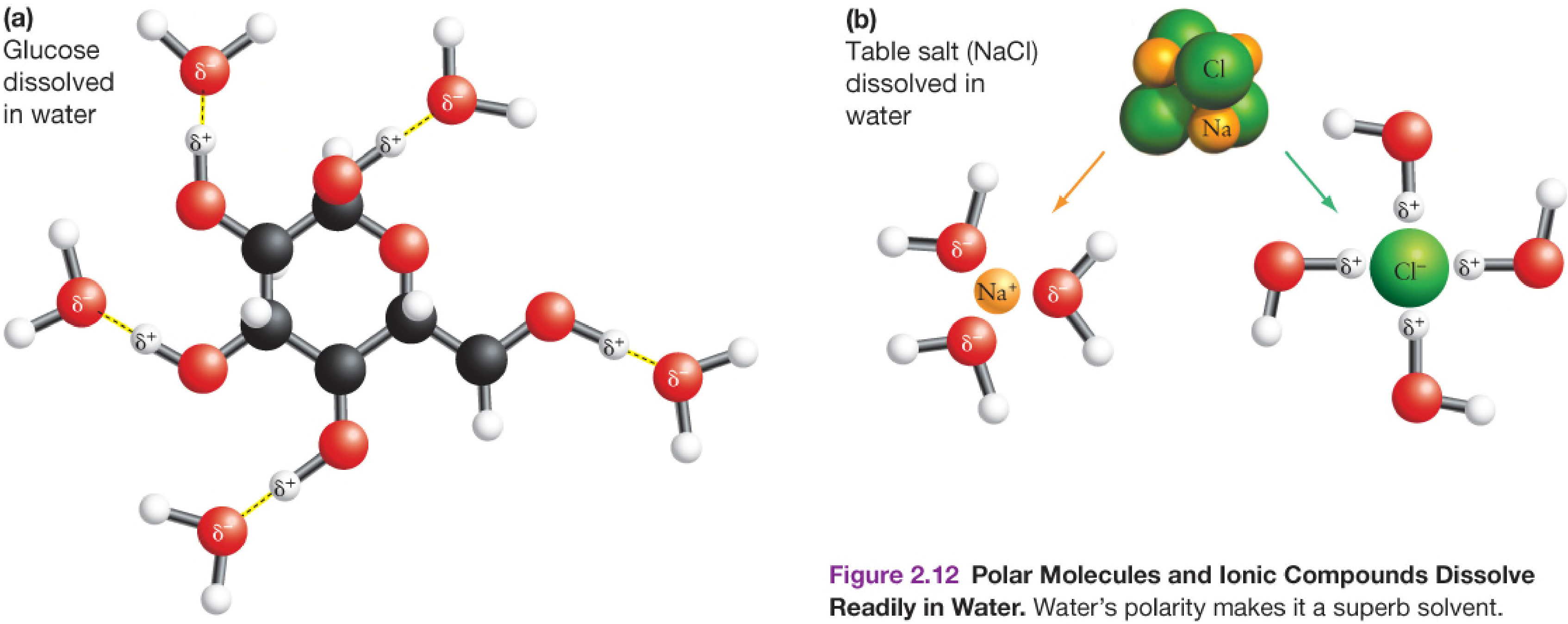


Figure 2.12 Polar Molecules and Ionic Compounds Dissolve Readily in Water. Water's polarity makes it a superb solvent.

Why Is Water Such an Efficient Solvent?

- **Hydrophobic** ("water-fearing") molecules
  - Are uncharged and nonpolar compounds
  - They do not dissolve in water
- Hydrophobic molecules interact with each other through **hydrophobic interactions**

Figure 2.13 Nonpolar Molecules Do Not Dissolve in Water

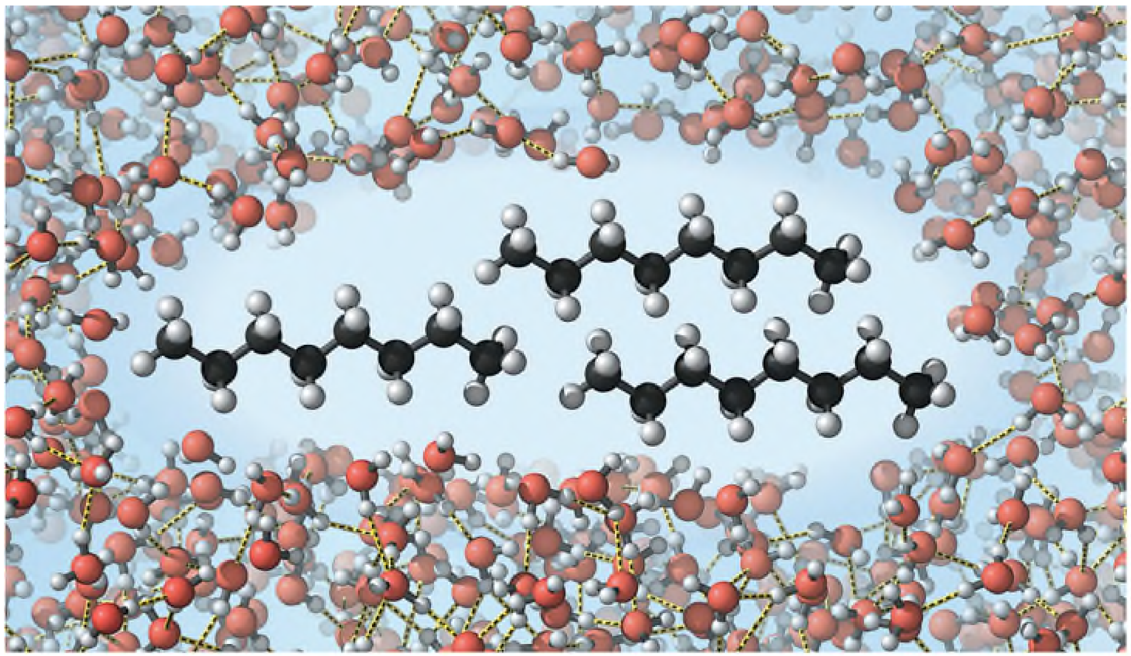


Figure 2.13 Nonpolar Molecules Do Not Dissolve in Water. In aqueous solution, nonpolar molecules such as octane (C_8H_{18})—a component of gasoline—are forced to interact with themselves. This occurs because water is much more stable when it interacts with itself rather than with the nonpolar molecules.

What Properties Are Correlated with Water's Structure?

- Water is unique due to its structure
  - Small size
  - Bent shape
  - Highly polar covalent bonds
  - Overall polarity

Cohesion, Adhesion, and Surface Tension

- Water also has several remarkable properties, largely due to its ability to form hydrogen bonds:
  1. Cohesive
  2. Adhesive
  3. Denser as a liquid than a solid
  4. Able to absorb large amounts of energy

Cohesion, Adhesion, and Surface Tension
(2 of 2)

- **Cohesion** – binding between like molecules
  - Water binds to itself by hydrogen bonding
  - Results in high **surface tension**
- **Adhesion** is binding between unlike molecules
  - Water binds to plastic or glass
  - Results in capillary action and meniscus formation

Figure 2.14 Cohesion, Adhesion, and Surface Tension

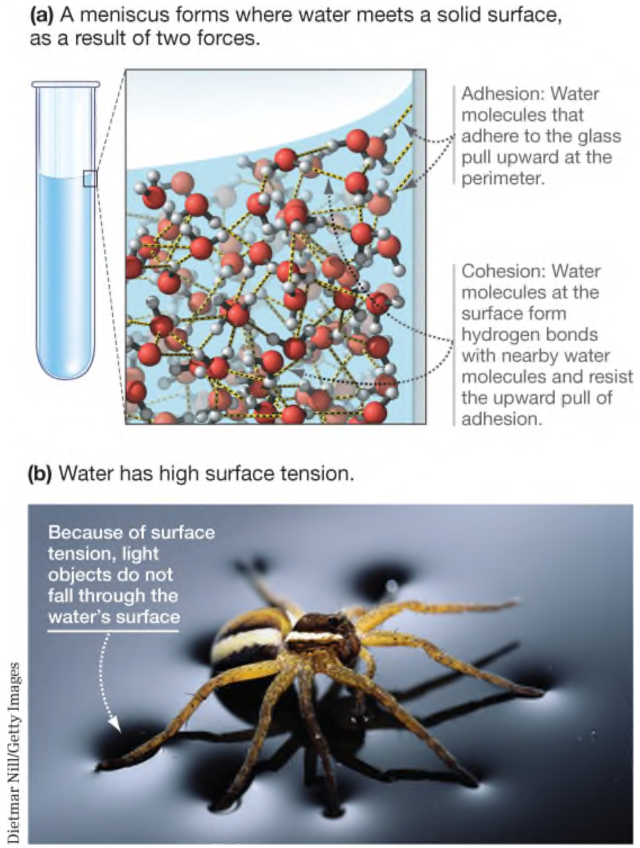


Figure 2.14 Cohesion, Adhesion, and Surface Tension.
(a) Meniscus formation is based on hydrogen bonding and other interactions with glass that are represented here by highlighted dashed lines. **(b)** Water resists forces—like the weight of a spider—that increase its surface area. The resistance is great enough that light objects do not break the surface.

Water Is Denser as a Liquid than as a Solid

- Most substances shrink as they solidify
- Water expands as it freezes
  - It is denser as a liquid than a solid
  - Forms a relatively open crystal structure
  - This is why ice floats!
  - Ice forms an insulating “blanket” on water surfaces

Figure 2.15 Hydrogen Bonding in Ice and Water

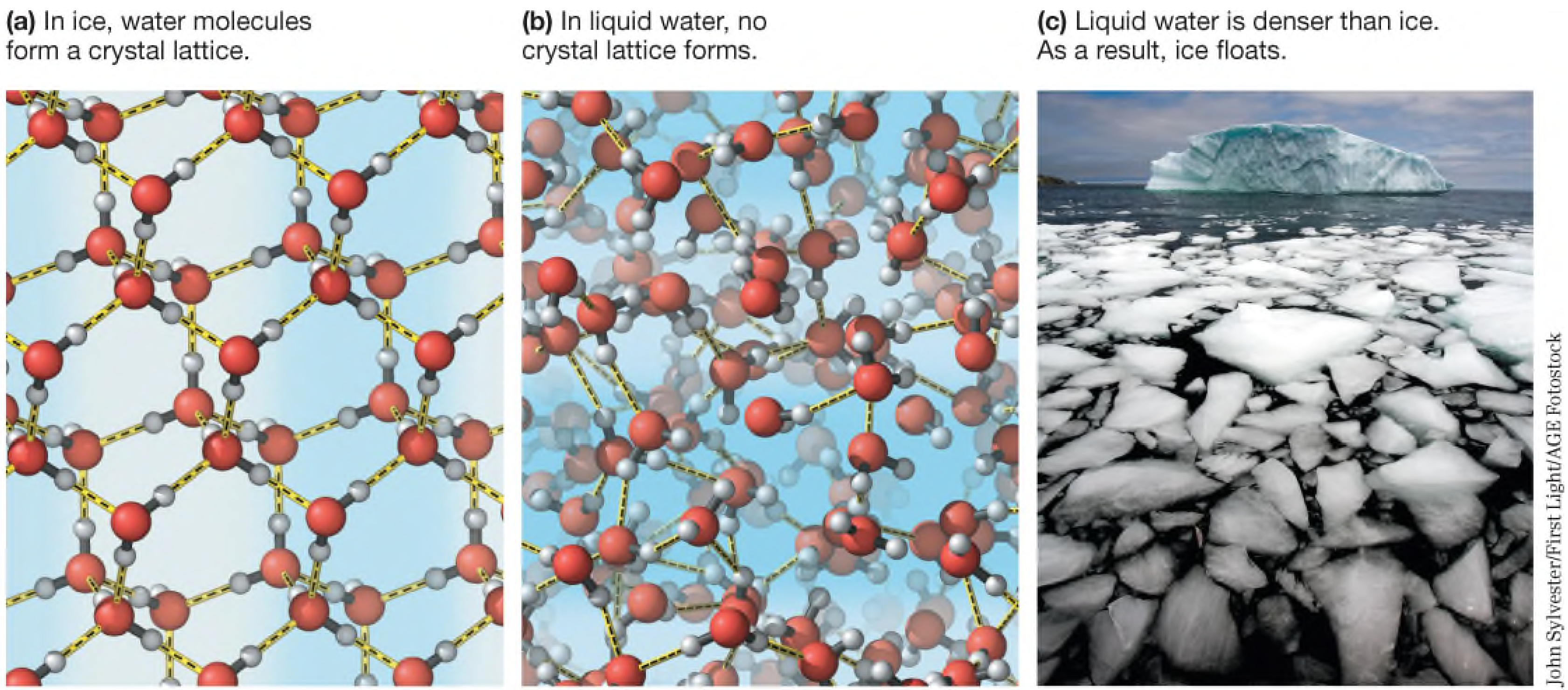


Figure 2.15 Hydrogen Bonding in Ice and Water. (a) In ice, each molecule forms four hydrogen bonds (yellow dashed lines) at one time. **(b)** As a liquid, bonds are continually broken and formed, so no lattice develops. **(c)** As a result, ice is less dense than water.

The Role of Water in Acid–Base Chemical Reactions (1 of 3)

- **Chemical reactions** occur when a substance is
  - Combined with another
  - Or broken down into another substance
- In most chemical reactions, chemical bonds are broken and new bonds form
- Chemical reactions are written as equations:
  - $\textbf{Reactant(s)} \, f \quad \textbf{Product(s)}$
  - For example: $H_2O \, f \quad H^+ + OH^-$

The Role of Water in Acid–Base Chemical Reactions (2 of 3)

- Water molecules dissociate into a **hydrogen ion** (H^+) and a **hydroxide ion** (OH^-)

$$H_2O\, f \quad H^+ + OH^-$$

- This happens in both directions at approximately the same rate = **chemical equilibrium**
- Since protons (H^+) don't exist by themselves, the reaction actually produces hydronium ions (H_3O^+)

$$2H_2O\, f \quad H_3O^+ + OH^-$$

The Role of Water in Acid–Base Chemical Reactions (3 of 3)

- **Acids** are substances that give up protons during chemical reactions and raise the hydronium ion concentration $[H_3O^+]$
  - Adding an acid to a solution increases the proton concentration of the solution
- **Bases** are substances that acquire protons during chemical reactions and lower $[H_3O^+]$
  - Adding a base to a solution decreases the proton concentration

Determining the Concentration of Protons

- The **molecular weight** of a molecule is the sum of the atomic weights of all the atoms in the molecule
- One **mole**
  - Equals 6.022×10^{23} molecules
  - Has a mass equal to the molecular weight expressed in grams
- The concentration of a substance in a solution is typically expressed as **molarity** (M)
  - Molarity is the number of moles per litre

The pH of a Solution Reveals Whether It Is Acidic or Basic (1 of 2)

- The number of protons in a solution determines how acid–base reactions occur
- There is no simple way to count protons
  - The concentration in water is very low
  - 1×10^{-7} M
- The **pH** scale expresses proton concentration $[H^+]$ in a solution
  - Negative base 10 logarithmic scale
  - pH of water is 7

The pH of a Solution Reveals Whether It Is Acidic or Basic (2 of 2)

- Acids have a pH of less than 7
- Bases have a pH of greater than 7
- **Buffers** protect against changes in pH
  - Life is sensitive to pH
  - Buffers help maintain **homeostasis** = relatively constant conditions

Figure 2.16 The pH Scale

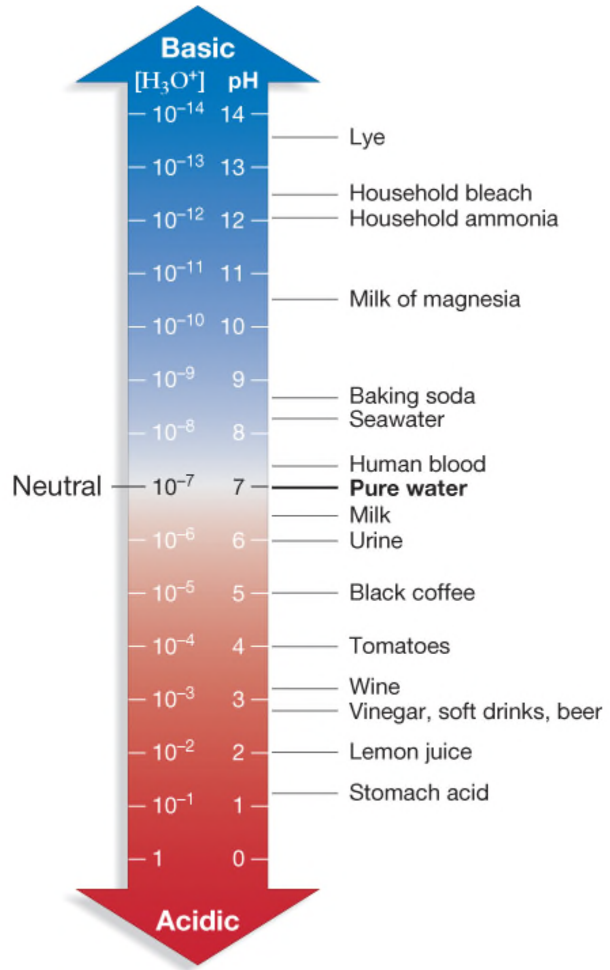


Figure 2.16 The pH Scale. Because the pH scale is logarithmic, a change in one unit of pH represents a change in the concentration of hydrogen ions equal to a factor of 10. Coffee has a hundred times more H^+ than pure water has.

Buffers Protect Against Damaging Changes in pH

- Life is sensitive to changes in pH
- Changes in proton concentration affect the structure and function of polar or charged substances
  - As well as the tendency of acid–base reactions to occur
- **Buffers** minimize changes in pH
  - Maintain **homeostasis** - relatively constant conditions
  - Carbonic acid works as a buffer in blood
    - disassociation of carbonic acid in an aqueous solution to form bicarbonate ions and protons

Chemical Reactions, Energy, and Chemical Evolution

- Chemical evolution may have begun in:
  1. *The atmosphere*, which was probably dominated by volcanic gases:
     - Mostly water vapor, carbon dioxide (CO_2), and nitrogen (N_2)
  2. *Deep-sea hydrothermal vents*, which have
     - Extremely hot rocks
     - Gases such as CO_2 and H_2
     - Minerals with reactive metals

How Do Chemical Reactions Happen? (1 of 2)

- The most common reaction in mix of gases from volcanoes produces carbonic acid:

$$CO_2(g) + H_2O(g) \rightleftharpoons CH_2O_3(aq)$$

- This expression is balanced
  - Same number of atoms on each side
- Equilibrium can be disturbed by
  - Changing the concentration of reactants or products
  - Changing the temperature

How Do Chemical Reactions Happen? (2 of 2)

- A **system** is a set of interacting components
- For example:

$$H_2O(l) \rightleftharpoons H_2O(g)$$

- If this system absorbs enough thermal energy from the environment
  - Liquid water $H_2O(l)$ will convert to gas $H_2O(g)$
- **Endothermic** reactions must absorb thermal energy to proceed
- **Exothermic** reactions release thermal energy

What Is Energy? (1 of 4)

- **Energy** is the capacity to do work or supply heat
- This capacity exists in one of two ways:
  1. **Potential energy—**stored potential to do work
  2. **Kinetic energy—**active energy of movement

What Is Energy? (2 of 4)

- In molecules, potential energy is related to the position of shared electrons in covalent bonds
  - If the shared electrons are far from the atoms' nuclei, the bonds are long and weak
  - If the electrons are shifted closer to one or both nuclei, the bond becomes shorter and stronger
- A molecule's potential to form stronger bonds is a type of potential energy called **chemical energy**

Figure 2.17 Potential Energy as a Function of Electron Sharing

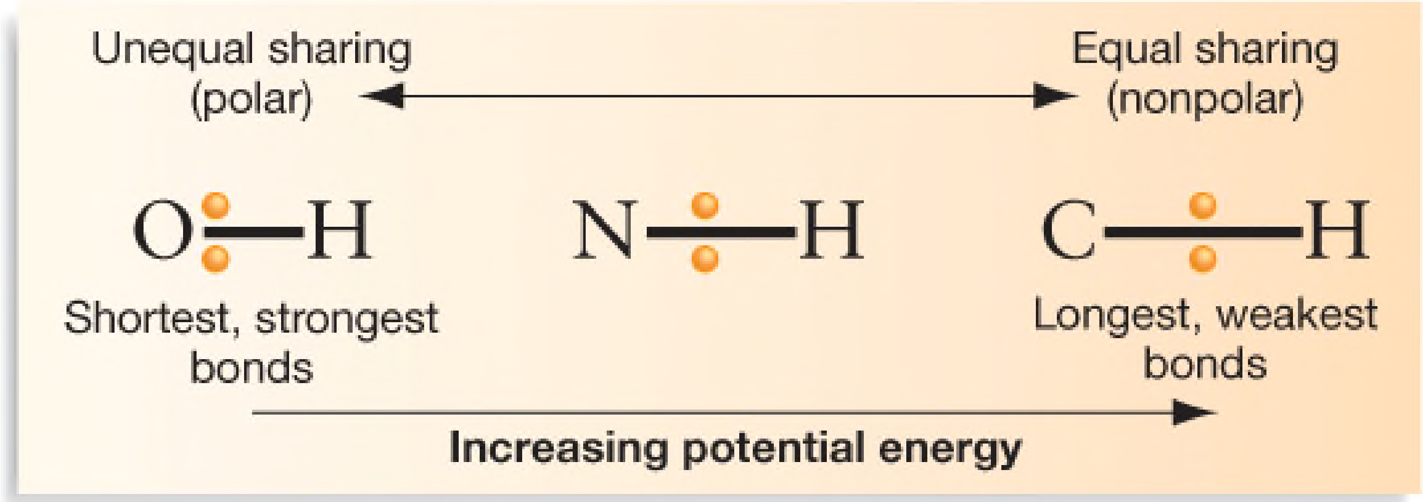


Figure 2.17 Potential Energy as a Function of Electron Sharing. Highly electronegative atoms, such as oxygen, pull shared electrons closer to their own nuclei, increasing bond strength and decreasing the potential energy of a molecule. Less electronegative atoms, such as carbon and hydrogen, share electrons more equally, decreasing bond strength and increasing potential energy.

What Is Energy? (3 of 4)

- The kinetic energy of molecular motion is called **thermal energy**
  - Molecules are constantly in motion
  - **Temperature** is a measure of the thermal energy in a molecule
    - If an object has a low temperature, its molecules are moving slowly; we perceive this as “cold”
    - If an object has a high temperature, its molecules are moving rapidly; we perceive this as “hot”
  - **Heat** is a measure of thermal energy being transferred between two objects

What Is Energy? (4 of 4)

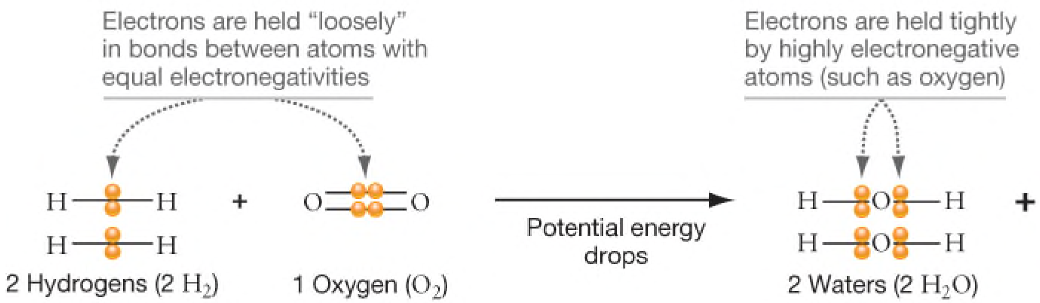
- The **first law of thermodynamics**
  - Energy is conserved
  - It cannot be created or destroyed
  - It can be transferred or transformed
- Energy transformation is the heart of chemical evolution
  - Molecules of the early Earth were exposed to massive energy input

What Makes a Chemical Reaction Spontaneous? (1 of 2)

- Chemical reactions are spontaneous if
  - They proceed without any continuous external influence
  - No added energy is needed
- Spontaneity of a reaction is determined by two factors:
  1. Products are less ordered than the reactants
     - **Entropy** (disorder) increases
  2. Products have lower potential energy than the reactants (shared electrons are held more tightly in the reactants)

Figure 2.18 Potential Energy May Change during Chemical Reactions

(a) When hydrogen and oxygen gas react, the products have much lower potential energy than the reactants.



(b) The difference in potential energy is released as heat and light, which vaporizes the water produced.

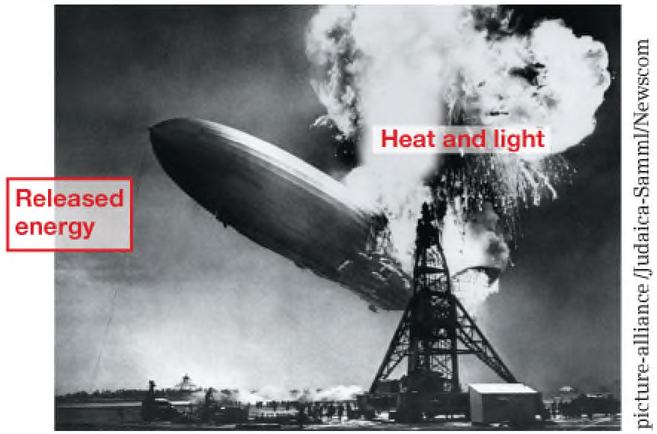


Figure 2.18 Potential Energy May Change during Chemical Reactions. In the Hindenburg disaster of 1937, hydrogen gas from a lighter-than-air craft reacted with oxygen in the atmosphere, with devastating results.

What Makes a Chemical Reaction Spontaneous? (2 of 2)

- The **second law of thermodynamics**
  - Entropy always increases
  - Chemical reactions result in products with
    - Less ordered energy
    - Less usable energy
- Physical and chemical processes proceed in the direction that results *in lower potential energy and increased disorder*

Figure 2.19 Spontaneous Processes Result in Lower Potential Energy, Increased Disorder, or Both

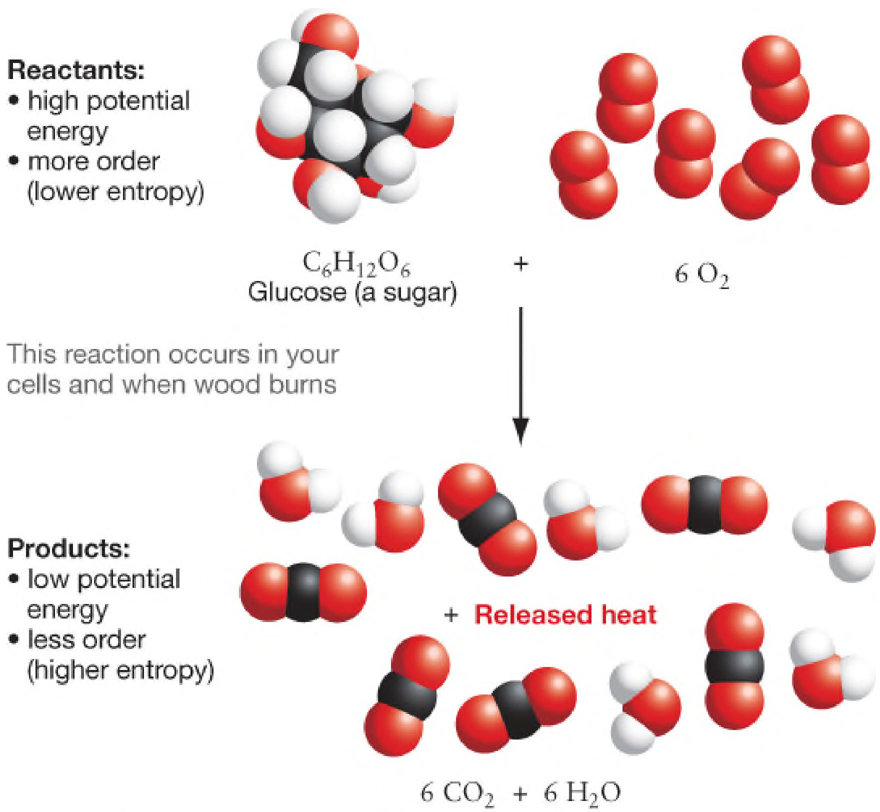


Figure 2.19 Spontaneous Processes Result in Lower Potential Energy, Increased Disorder, or Both.

Model Systems for Investigating Chemical Evolution

- Two systems that model chemical evolution:
  1. **Prebiotic soup model**
     - Certain molecules were synthesized from gases in the atmosphere or arrived via meteorites
     - Condensed with rain and accumulated in oceans
     - Result in an "organic soup" that allowed for continued construction of larger, even more complex molecules
  2. **Surface metabolism model**
     - Dissolved gases came in contact with minerals lining the walls of deep-sea vents
     - Formed more complex, organic molecules

Early Origin-of-Life Experiments

- Stanley Miller wanted to answer a simple question:
  - Can complex organic compounds be synthesized from the simple molecules present in Earth's early atmosphere?
  - Put another way, is it possible to re-create the first steps in chemical evolution by simulating early-Earth conditions in the laboratory?

Figure 2.20 Miller’s Spark-Discharge Experiment

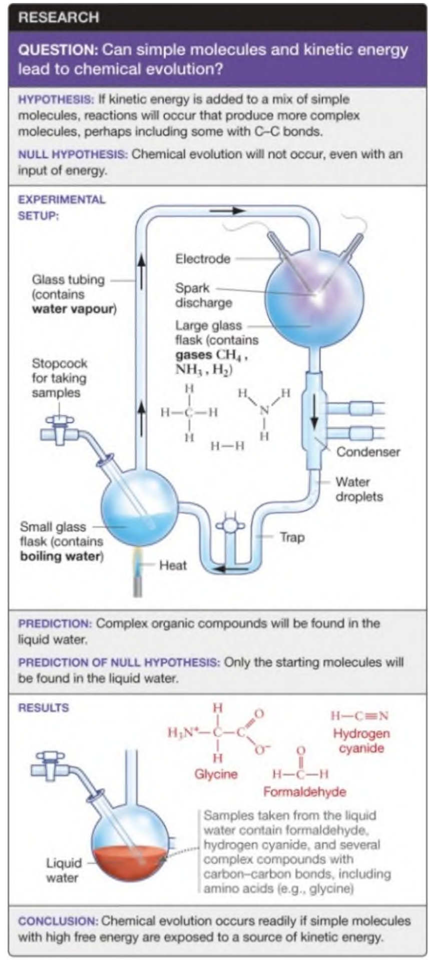


Figure 2.20 Miller’s Spark-Discharge Experiment. The arrows in the “Experimental Setup” diagram indicate the flow of water vapour or liquid. The condenser is a jacket with cold water flowing through it.

SOURCE: Based on Miller, S. L. 1953. A production of amino acids under possible primitive Earth conditions. Science 117: 528–529.

Recent Origin-of-Life Experiments

- Miller's apparatus showed complex molecules could be formed from simple molecules
  - Used heat and electrical charges
  - Formed precursors to life molecules
- Later experiments synthesized precursors using light energy in the form of high-energy **photons**
  - Created highly reactive **free radicals**

Figure 2.21 Free Radicals Are Extremely Reactive

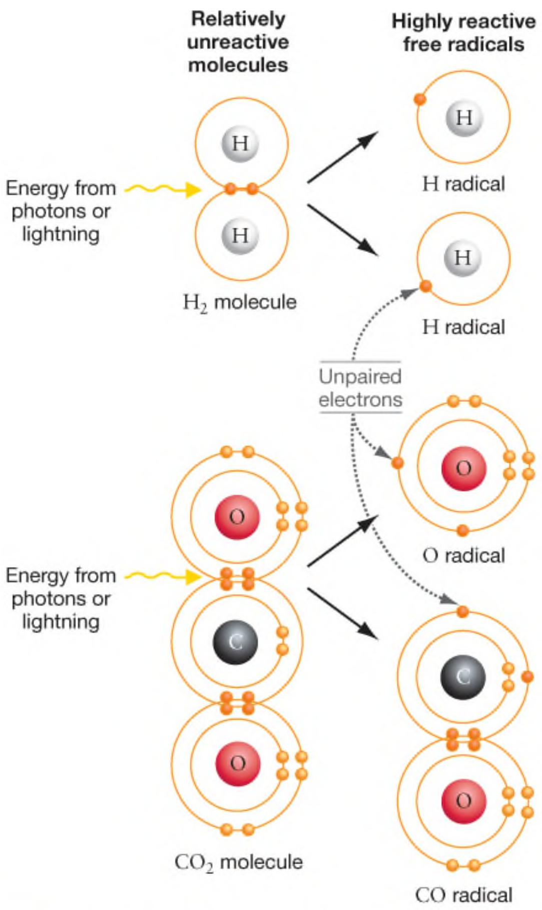
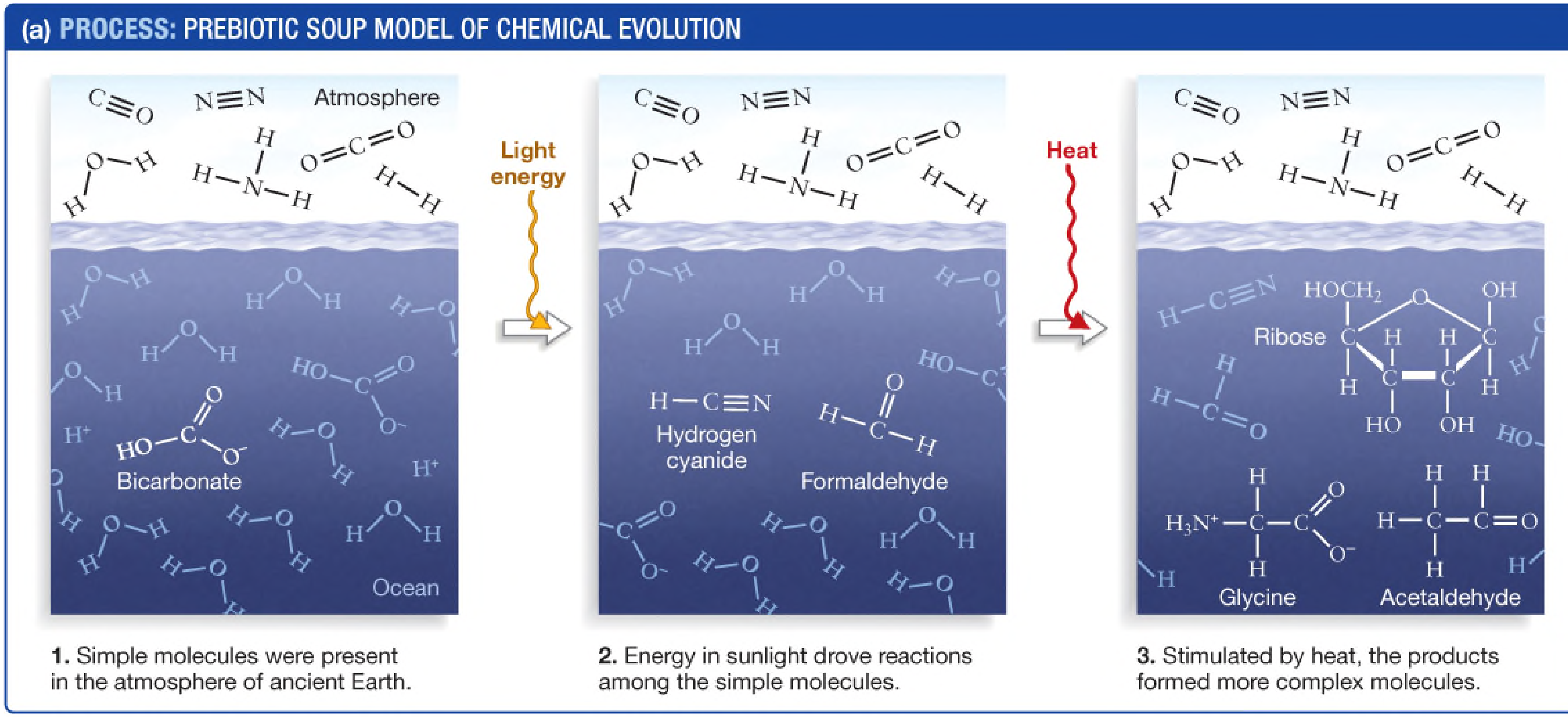


Figure 2.21 Free Radicals Are Extremely Reactive. When high-energy photons or the electrical energy from lightning strike molecules of hydrogen or carbon dioxide, free radicals can be created. Formation of free radicals is thought to be responsible for some key reactions in chemical evolution.

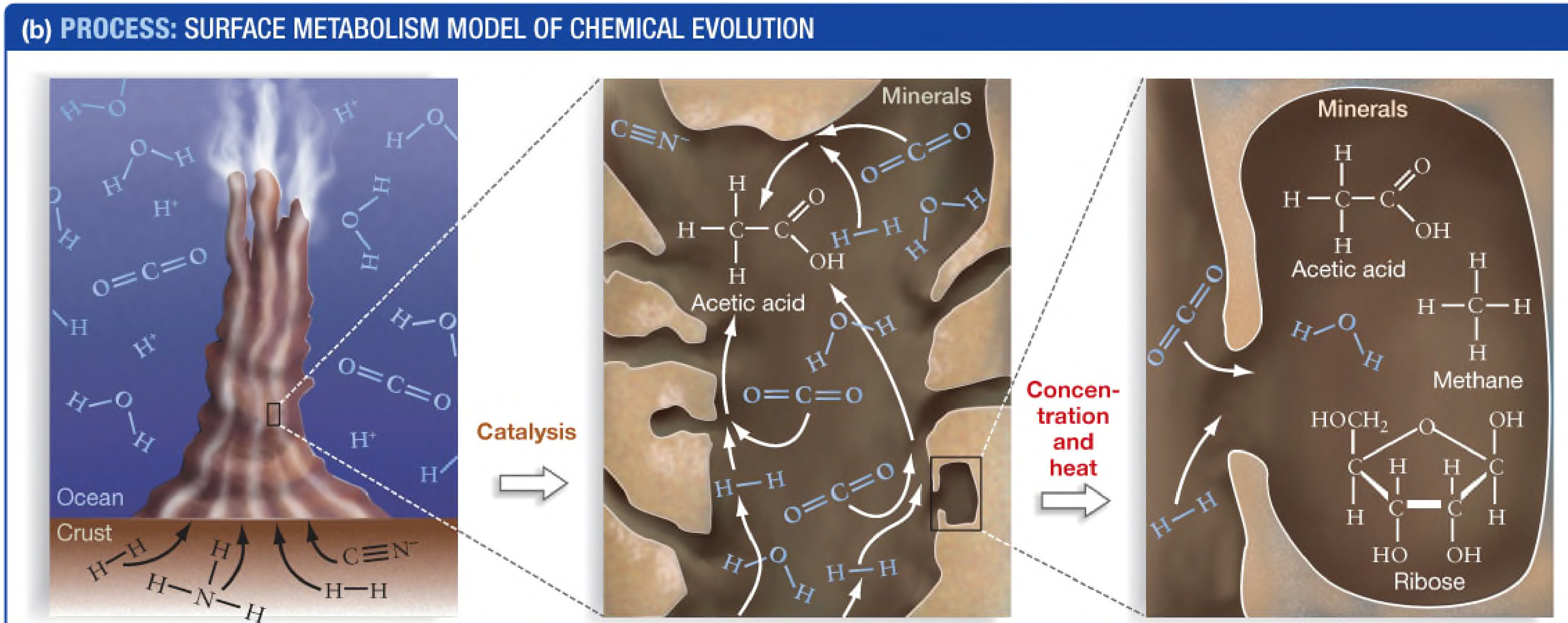
Figure 2.22 The Start of Chemical Evolution—Two Models (1 of 2)



Synthesis of Precursors Using Light Energy

- Reactive molecules must have been localized and concentrated once they entered the ocean
- Surface metabolism model proposes that reactants were attracted to mineral deposits in deep-sea vents
- Minerals would also act as **catalysts** to speed up chemical reactions in the vents
- Preliminary research shows that large carbon-based molecules can be formed under these conditions

Figure 2.22 The Start of Chemical Evolution—Two Models (2 of 2)



1. Simple molecules were present in early oceans and hydrothermal vents.

2. Vent minerals catalyzed spontaneous reactions among high-energy molecules.

3. Stimulated by heat and concentration, the products formed more complex molecules.

Concentration and Catalysis in Hydrothermal Vents

- A major stumbling block in the prebiotic soup model is
  - Precursor molecules would have become diluted when they entered the early oceans
- The surface metabolism model
  - reactants are recruited to a defined space
  - a layer of reactive minerals deposited on the walls of deep-sea vent chimneys
- Did vent minerals serve as **catalysts** in the synthesis of acetic acid in early Earth?

Canadian Research 2.1: Searching for Life in Extreme Environments

- Researchers are looking for life in unusual places: **extremophiles** (extreme loving organisms)
- The Endeavour Hydrothermal Vents
  - Researchers found a rich community based upon
  - Simple organisms feeding on CO_2 and H_2
- Deep inside the Canadian Shield
  - Fluid was extracted from bore holes dated as 1.5 billions years old
  - Goal is to find out if the fluid could or does support life

The Importance of Organic Molecules

- Carbon is the most versatile atom on Earth
  - Because of its four valence electrons
  - Because it can form many covalent bonds
- **Organic compounds** are molecules that contain carbon bonded to other elements
  - An almost limitless array of molecular shapes
  - With different combinations of single and double bonds
- The formation of carbon–carbon bonds was an important event in chemical evolution

Figure 2.23 The Shapes of Carbon-Containing Molecules

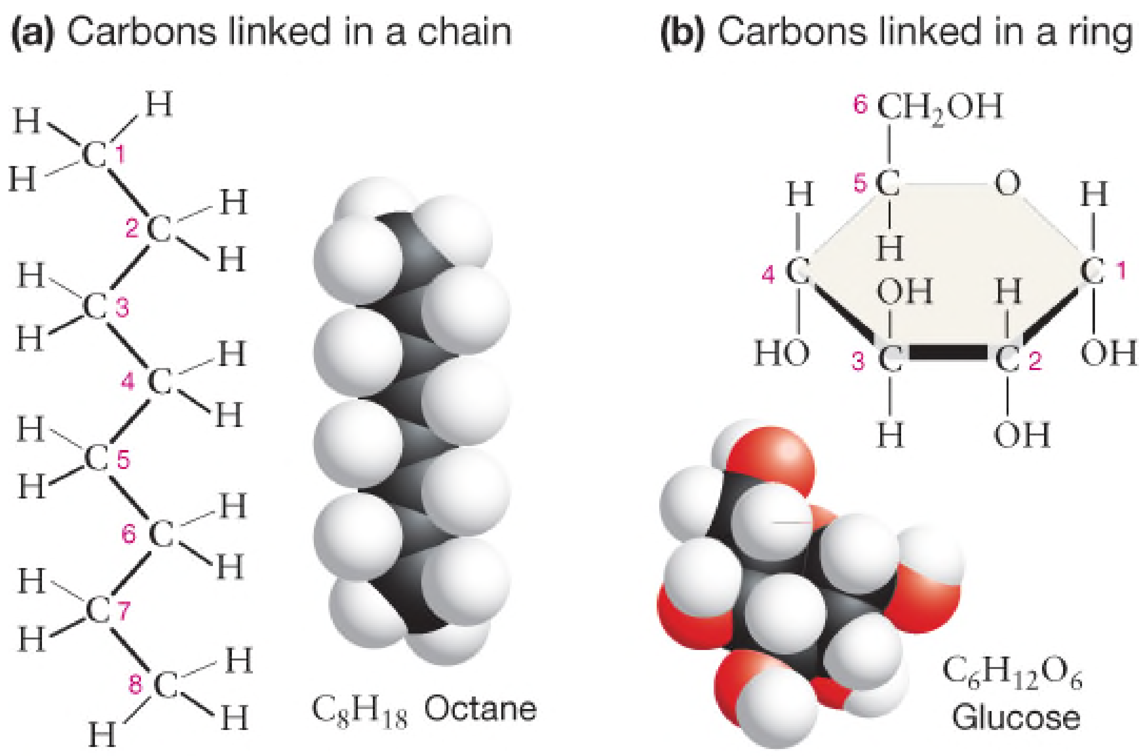


Figure 2.23 The Shapes of Carbon-Containing Molecules.
(a) Octane is a hydrocarbon chain present in gasoline. **(b)** Glucose is a sugar that can form a ring-like structure.

Canadian Research 2.1: The Carbon-Rich Tagish Lake Meteorite (1 of 3)

- Most asteroids land in oceans and unpopulated regions and are never found, but in January 2000 a meteorite landed near Atlin, BC.
- Fortunately when fragments were collected they were handled with gloves and were quickly frozen.
- This meteorite was unusual because:
  - It contained carbonaceous chondrites that date from the birth of our solar system.
  - It contained a lot of organic molecules, including amino acids.

Canadian Research 2.2: The Carbon-Rich Tagish Lake Meteorite (2 of 3)

- Carbon like most other atoms is created within older stars by thermonuclear reactions fusing helium nuclei together.
- The meteorite was analyzed by Peter Brown from the University of Western Ontario and others.
- **They found the asteroid contained:**
  - 3.7% carbonate minerals such as $FeCO_3$ and
  - 1.7% other types of carbon-containing molecules.

Canadian Research 2.2: The Carbon-Rich Tagish Lake Meteorite (3 of 3)

- Because the overall carbon content of this meteorite is
  - Unexpectedly high
  - This meteorite may be very old.
- The Tagish Lake meteorite may be the most primitive object ever studied in a laboratory.

Functional Groups

- Important H-, N-, O-, P-, and S-containing **functional groups** found in organic compounds:
  - **Amino groups** attract a proton and act as bases
  - **Carboxyl** drop a proton, and act as acids
  - **Carbonyl groups have sites** that link molecules into larger, more-complex compounds
  - **Hydroxyl groups a**ct as weak acids
  - **Phosphate groups** have two negative charges
  - **Sulfhydryl groups** link together via disulfide bonds

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Table 2.1 Six Functional Groups Commonly Attached to Carbon Atoms

SUMMARY **Table 2.1** Six Functional Groups Commonly Attached to Carbon Atoms

Functional Group	Formula*	Family of Molecules	Properties of Functional Group	Example
Amino	H_2N—R	Amines	Acts as a base—tends to attract a proton to form: H_3N^+—R	H_2N—CH_2—COOH Glycine (an amino acid)
Carboxyl	R—C(=O)OH	Carboxylic acids	Acts as an acid—tends to lose a proton in solution to form: R—C(=O)O^-	CH_3—C(=O)OH Acetic acid
Carbonyl	R—C(=O)H	Aldehydes	Aldehydes, especially, react with certain compounds to produce larger molecules to form: R1—C(=O)—H + H—R2 → R1—C(OH)(H)—R2 (R group from aldehyde; R group from another reactant)	CH_3—C(=O)H Acetaldehyde
	R—C(=O)—R	Ketones		CH_3—C(=O)—CH_3 Acetone
Hydroxyl	R—OH	Alcohols	Highly polar, so makes compounds more soluble through hydrogen bonding with water; may also act as a weak acid and drop a proton	CH_3—CH_2—OH Ethanol
Phosphate	R—O—P(=O)(O^-)—O^-	Organic phosphates	Molecules with more than one phosphate linked together store large amounts of chemical energy	HO—C(=O)—CH(OH)—CH_2—O—P(=O)(O^-)—O^- 3-Phosphoglyceric acid
Sulfhydryl	R—SH	Thiols	When present in proteins, can form disulfide (S–S) bonds that contribute to protein structure	H_3N^+—CH(CH_2—SH)—C(=O)O^- Cysteine

*In these structural formulas, "R" stands for the rest of the molecule.

Learning Objectives: *You should be able to …*

- Describe how and why atoms interact to form molecules. Sketch examples of how electron pairs are shared in nonpolar covalent bonds, polar covalent bonds, and ionic bonds.
- List the unique properties of water. Explain how these properties relate to the structure of water molecules.
- Explain how the structure of water explains its biologically important properties.
- Define energy and describe some of the major forms that energy can take. Explain why chemical bonds can be considered a form of potential energy.

Learning Objectives: *You should be able to …*
(2 of 2)

- Explain in simple terms how changes in entropy and potential energy determine whether or not a reaction is spontaneous.
- Describe the current models for chemical evolution on the early Earth.
- Explain why carbon is a key element for life on Earth. List the six major functional groups, their structural formulas, and their basic characteristics.