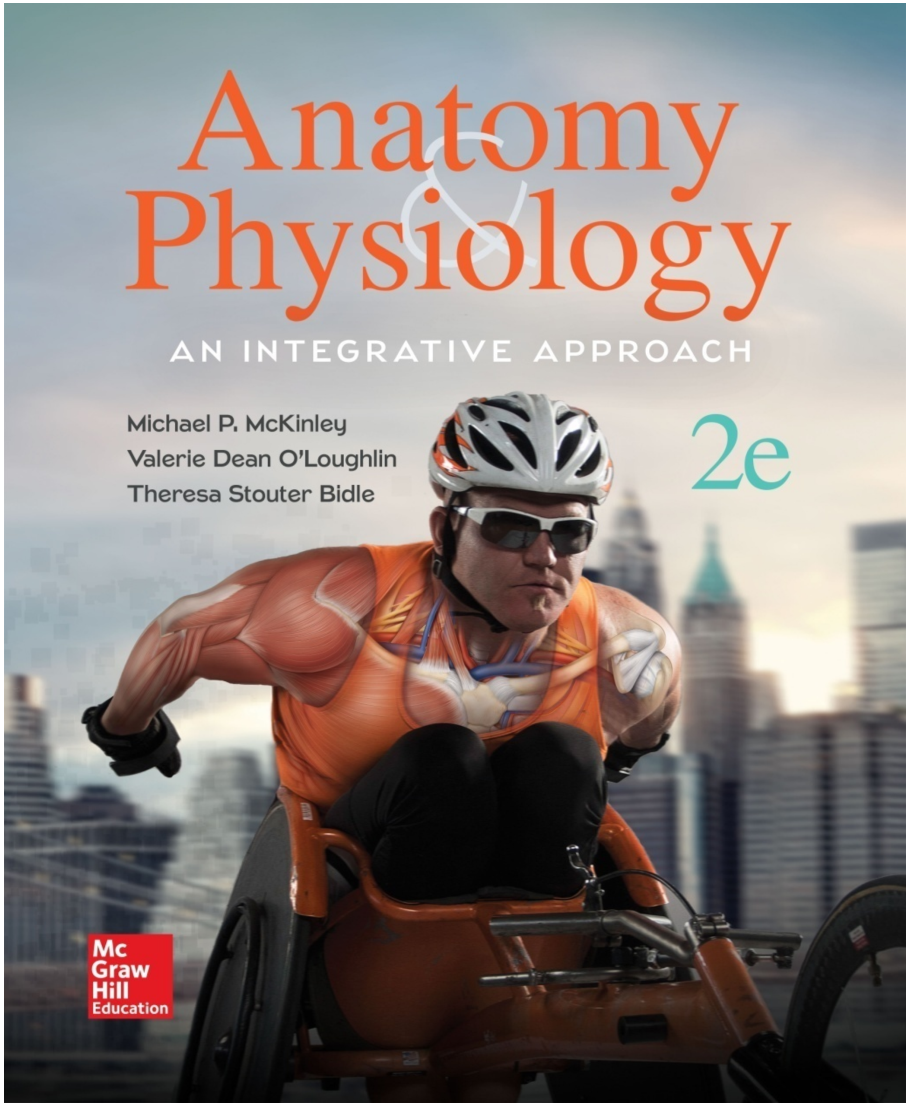
Chapter 02
Lecture Outline

See separate PowerPoint slides for all figures and tables pre-inserted into PowerPoint without notes.



2.1 Atomic Structure

Learning Objectives:

1. Define matter, and list its three forms.
2. Describe and differentiate among the subatomic particles that compose atoms.
3. Explain the arrangement of elements in the periodic table based on atomic number.
4. Diagram the structure of an atom.

2.1 Atomic Structure (*continued*)

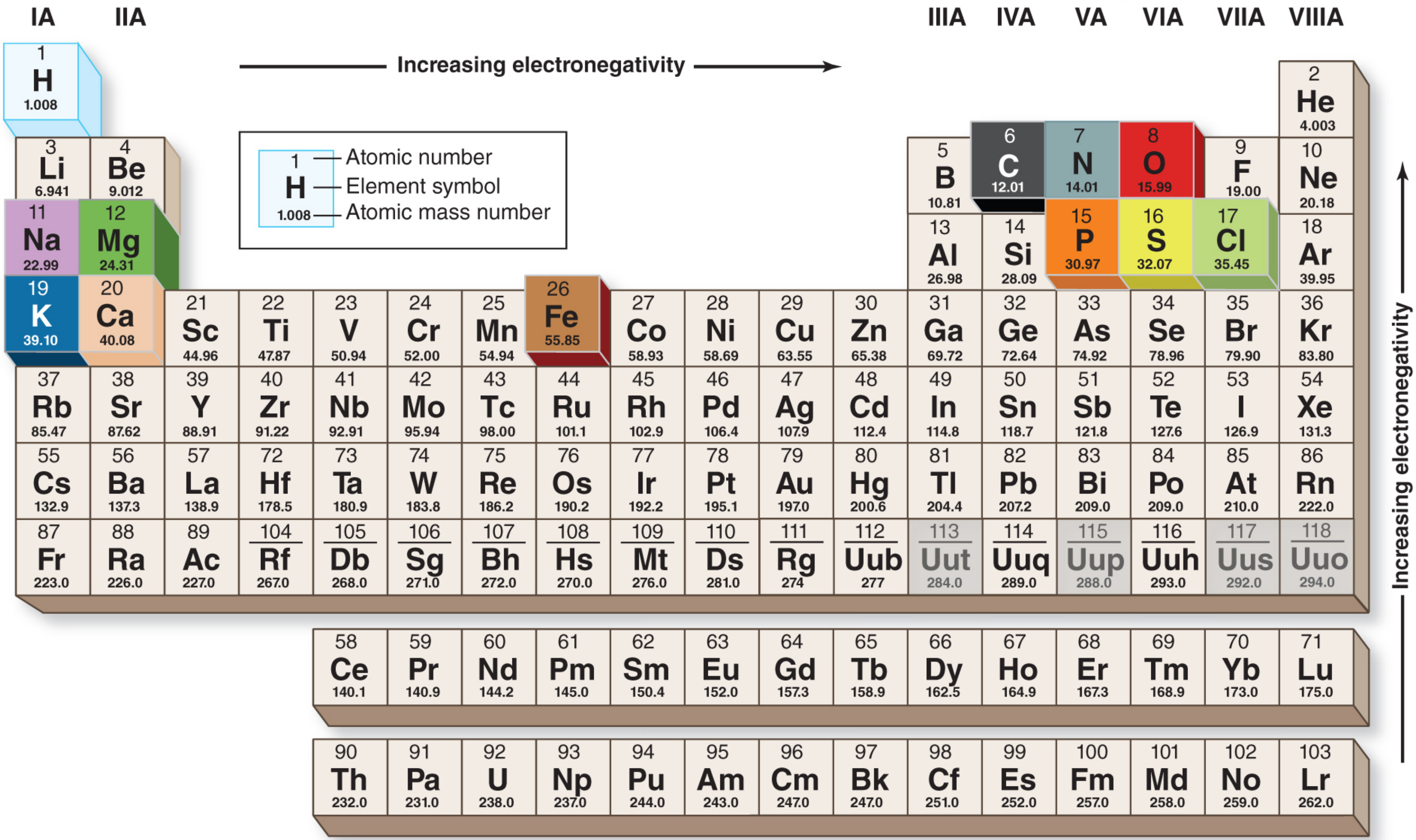
Learning Objectives:

5. Describe an isotope.
6. Explain how radioisotopes differ from other types of isotopes.
7. Describe how elements are organized in the periodic table based on the valence electron number.
8. State the octet rule.

2.1a Matter, Atoms, Elements, and the Periodic Table

- **Matter** has mass and occupies space
  - Composes human body
  - 92 naturally occurring **elements** make up matter
  - **Atom** is the smallest particle exhibiting chemical properties of an element
  - 3 forms of matter
    - Solid (e.g., bone)
    - Liquid (e.g., blood)
    - Gas (e.g., oxygen)

Periodic Table of Elements



(a)

Figure 2.1a

Most Common Elements of the Human Body

Major elements
(collectively compose more than 98% of body weight)

Symbol		% Body weight
O	Oxygen	65.0
C	Carbon	18.0
H	Hydrogen	10.0
N	Nitrogen	3.0
Ca	Calcium	1.5
P	Phosphorus	1.0

Minor elements
(collectively compose less than 1% of body weight)

Symbol		% Body weight
S	Sulfur	0.25
K	Potassium	0.20
Na	Sodium	0.15
Cl	Chlorine	0.15
Mg	Magnesium	0.05
Fe	Iron	0.006

(b)

Figure 2.1b

2.1a Matter, Atoms, Elements, and the Periodic Table

Components of an atom

- Atoms composed of three **subatomic particles**
  - **Neutrons**
    - Mass of one **atomic mass unit (amu)**
    - No charge
  - **Protons**
    - Mass of one amu
    - Positive charge of one (+1)
  - **Electrons**
    - 1/1800th of the mass of a proton or neutron
    - Negative charge of one (–1)
    - Located at varying distance from the nucleus in regions called *orbitals*

2.1a Matter, Atoms, Elements, and the Periodic Table

- Periodic table
  - **Chemical symbol**
    - Unique to each element
    - Usually identified by first letter, or first letter plus an additional letter, e.g., C is carbon
  - **Atomic number**
    - Number of protons in an atom of the element
    - Located above symbol name
    - Elements arranged by anatomic number within rows
  - **Average atomic mass**
    - Mass of both protons and neutrons
    - Shown below the element’s symbol on the table

2.1a Matter, Atoms, Elements, and the Periodic Table

Determining the number of subatomic particles

- Proton number = atomic number
- Neutron number = atomic mass – atomic number
    - Neutron number = (p + n) – p
    - Neutron number of Na = 23 – 11 = 12
- Electron number = proton number

2.1a Matter, Atoms, Elements, and the Periodic Table

Diagramming atomic structures

- An atom has shells of electrons surrounding the nucleus
  - Each shell has given energy level
  - Each shell holds a limited number of electrons
  - Innermost shell—two electrons, second shell up to eight
  - Shells close to the nucleus must be filled first

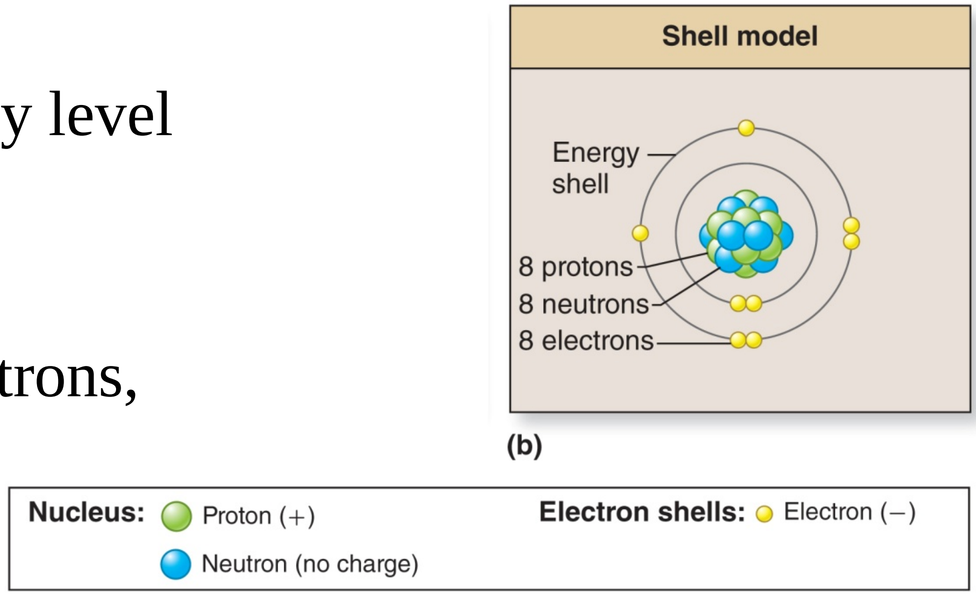


Figure 2.2b

What did you learn?

- What subatomic particles determine the mass of an atom?
- What subatomic particles determine the charge of an atom?

2.1b Isotopes

- **Isotopes** are different atoms of the same element
  - Same number of protons and electrons; different number of neutrons
  - Identical chemical characteristics; different atomic masses
- E.g., carbon exists in three isotopes
  - Carbon-12, with 6 neutrons
    - Most prevalent type
  - Carbon-13, with 7 neutrons
  - Carbon-14, with 8 neutrons

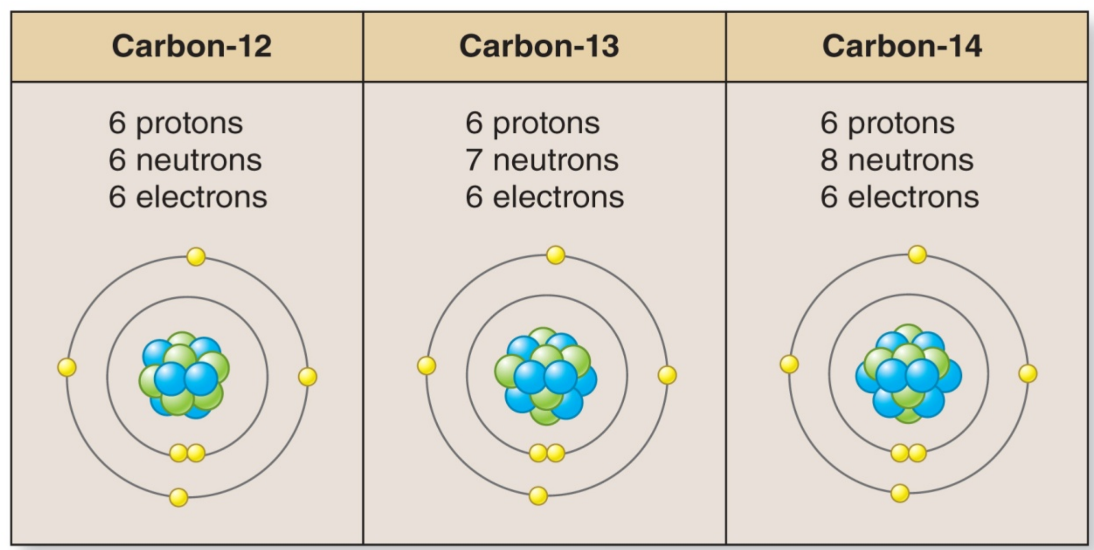


Figure 2.3

2.1b Isotopes

- Weighted average of atomic mass for all isotopes is the average atomic mass
- **Radioisotopes**
  - Contain excess neutrons, so unstable
  - Lose nuclear components in the form of high energy radiation
    - Alpha particles
    - Beta particles
    - Gamma rays

2.1b Isotopes

- **Physical half-life**
  - The time for 50% of radioisotope to become stable
  - Can vary from a few hours to thousands of years
- **Biological half-life**
  - The time required for half of the radioactive material from a test to be eliminated from the body

Clinical View: Medical Imaging of the Thyroid Gland Using Iodine Radioisotopes

- Radioisotopes introduced into the body during medical procedures
  - Used by cells in a similar manner to nonradioisotopes
  - Can trace products of metabolic reactions that use these elements
  - Thyroid gland darker in areas where less radioactive iodine taken up
  - Can help locate a nodule

2.1c Chemical Stability and the Octet Rule

- Periodic table is organized into columns based on number of electrons in outer shell, referred to as the **valence shell**
  - Column IA shows hydrogen, lithium, sodium, potassium
    - All with one electron in their outer shell
  - Each consecutive column has one additional electron in outer shell
  - Elements in column VIIA each have a full valence shell
    - Results in chemical stability
    - Helium, neon, etc., chemically inert *noble gases*

Organization of the Periodic Table Based on Valence Shell Electrons

- Elements tend to lose, gain, or share electrons to obtain complete outer shells with eight electrons
  - Known as the **octet rule**

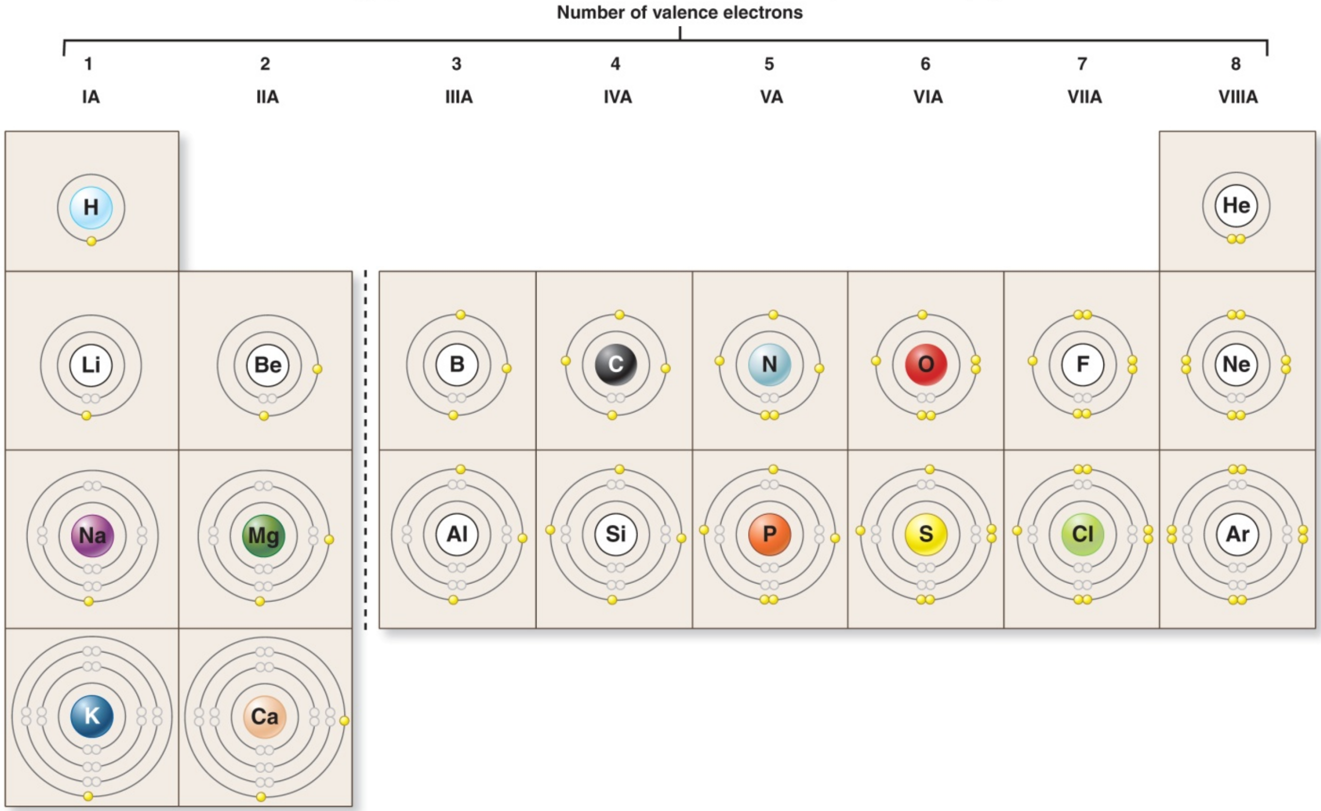


Figure 2.4

What did you learn?

- What is the relationship of the octet rule and chemical stability?

2.2 Ions and Ionic Compounds

Learning Objectives:

1. Define an ion.
2. List some common ions in the body.
3. Differentiate between cations and anions.
4. Describe how charges are assigned to ions.
5. Define an ionic bond.
6. Describe an ionic compound of NaCl.
7. List other examples of ionic compounds.

2.2 Ions and Ionic Compounds

- **Chemical compounds**
  - Stable associations between two or more elements combined in a fixed ratio
  - Classified as ionic or molecular
- Ionic compounds are structures composed of ions held together in a lattice by ionic bonds

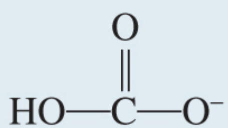
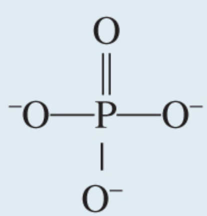
2.2a Ions

- **Ions**
  - Atoms with a positive or a negative charge
  - Produced from loss or gain of one or more electrons
  - Significant physiological functions
    - E.g., K^+ is used to sports drinks to replace the K^+ lost in sweat
    - E.g., K^+ in a large dose is used in some states for lethal injection

Common Ions in the Human Body and Their Physiological Significance (Table 2.1a)

COMMON CATIONS (POSITIVELY CHARGED IONS)		
Cation	**Structure**	**Physiologic Significance**
Sodium ion	Na^+	• Most common extracellular cation • Participant in conducting electrical signals in nerves and muscle • Most important in osmotic movement of water • Sodium gradient involved in cotransport of other substances across a plasma membrane
Potassium ion	K^+	• Most common intracellular cation • Participant in conducting electrical signals in nerves and muscle • Role in glycogen storage in liver and muscle • Function in pH balance
Calcium ion	Ca^{2+}	• Hardness of bone and teeth • Muscle contraction • Exocytosis (including release of neurotransmitter) • Blood clotting • Second messenger in hormonal stimulation of cells
Magnesium ion	Mg^{2+}	• Required for ATP production
Hydrogen ion	H^+	• Concentration determines pH of blood and other fluids of the body

COMMON ANIONS (NEGATIVELY CHARGED IONS)

Anion	Structure	Physiologic Significance
Chloride ion	Cl^-	• Alters nerve cell responsiveness to stimulation • Component of stomach acid (HCl) • Chloride shift in erythrocytes
Bicarbonate ion	HCO_3^- 	• Conversion of CO_2 gas to HCO_3^-, which is transported in the blood • Buffering of pH in blood
Phosphate ion	PO_4^{3-} 	• As $Ca_3(PO_4)_2$, it hardens bone and teeth • Component of phospholipids (membranes) • Component of nucleotides, including ATP and nucleic acids (DNA and RNA) • Most common intracellular anion • Intracellular buffer

Common Ions in the Human Body and Their Physiological Significance (Table 2.1b)

2.2a Ions

Losing electrons and the formation of cations

- Sodium can reach stability by donating an electron
  - Now satisfies the octet rule
  - Now has 11 protons and 10 electrons
  - Charge is +1
- **Cations** are ions with a positive charge

2.2a Ions

Gaining electrons and the formation of anions

- Chlorine reaches stability by gaining an electron
  - Now satisfies the octet rule
  - Now has 17 protons and 18 electrons
  - Charge is –1
- **Anions** are ions with negative charge
- **Polyatomic ions** are anions with more than one atom
  - E.g., bicarbonate ion and phosphate ion

2.2a Ions

General rules for assigning charges

- Atoms with 1, 2, or 3 electrons in valence shell become cations
  - E.g., calcium has two electrons in its outer shell
    - Reaches stability by donating two electrons
    - Develops a charge of +2
- Atoms with 5, 6, or 7 electrons become anions
  - E.g., chlorine has seven electrons in its outer shell
    - Reaches stability by gaining one electron
    - Develops a charge of –1

2.2a Ions

Using the periodic table to assign charges

- Elements on the left side ("metallic" side) tend to lose electrons
  - Specific charge dependent upon position of the element
    - Group 1A = +1, Group 2A = +2, Group 3A = +3
- Elements on the right side ("nonmetallic" side) tend to gain electrons
  - Specific charge dependent upon position of the element
    - Group VA = –3, Group VIA = –2, Group VIIA = –1

2.2b Ionic Bonds

- **Ionic bonds**
  - Cations and anions bound by electrostatic forces
  - Form **salts**
    - E.g., table salt (NaCl)
      - Each sodium atom loses one outer shell electron to a chlorine atom
      - Sodium and chlorine ions are held together by ionic bonds in a lattice crystal structure (**ionic compound**)
    - E.g., magnesium chloride
      - Each magnesium atom loses one electron to each of the two chlorine atoms

Formation of an Ionic Bond Involving Sodium and Chloride

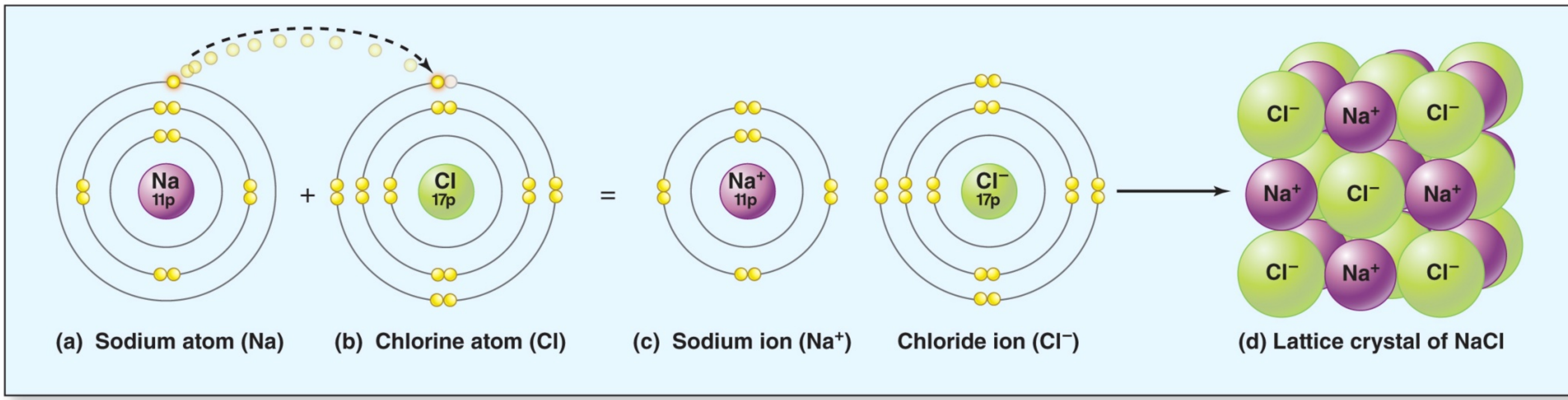


Figure 2.5

What did you learn?

- What is the charge for a magnesium ion, based on its position in the periodic table?

- Could an ionic bond form between two cations?

2.3 Covalent Bonding, Molecules, and Molecular Compounds

Learning Objectives:

1. Define a molecular formula.
2. Describe a structural formula, and explain its use in differentiating isomers.
3. Describe a covalent bond and explain its formation based on the octet rule.
4. List the four most common elements in the human body.
5. Distinguish between single, double, and triple covalent bonds.

2.3 Covalent Bonding, Molecules, and Molecular Compounds (*continued*)

Learning Objectives:

6. Explain polar and nonpolar covalent bonds.
7. Describe the difference between a nonpolar molecule and a polar molecule.
8. Define an amphipathic molecule.
9. Describe the hydrogen bonding between polar molecules.
10. List and define the intermolecular attractions between nonpolar molecules.

2.3 Covalent Bonding, Molecules, and Molecular Compounds

- **Covalently bonded molecule**
  - Electrons shared between atoms of two or more different elements
  - Termed **molecular compounds**
    - E.g., carbon dioxide (CO_2), but not molecular oxygen (O_2)

2.3a Chemical Formulas: Molecular and Structural

- **Molecular formula**
  - Indicates number and type of atoms
  - E.g., carbonic acid (H_2CO_3)
- **Structural formula**
  - Indicates number and type of atoms
  - Indicates arrangement of atoms within the molecule
  - E.g., O=C=O (carbon dioxide)
  - Allows differentiation of **isomers**
    - Same number and type of elements, but arranged differently in space

2.3a Chemical Formulas: Molecular and Structural

- Glucose vs. galactose vs. fructose
  - Same molecular formula
  - 6 carbon, 12 hydrogen, 6 oxygen
  - Atoms arranged differently
- Isomers may have different chemical properties

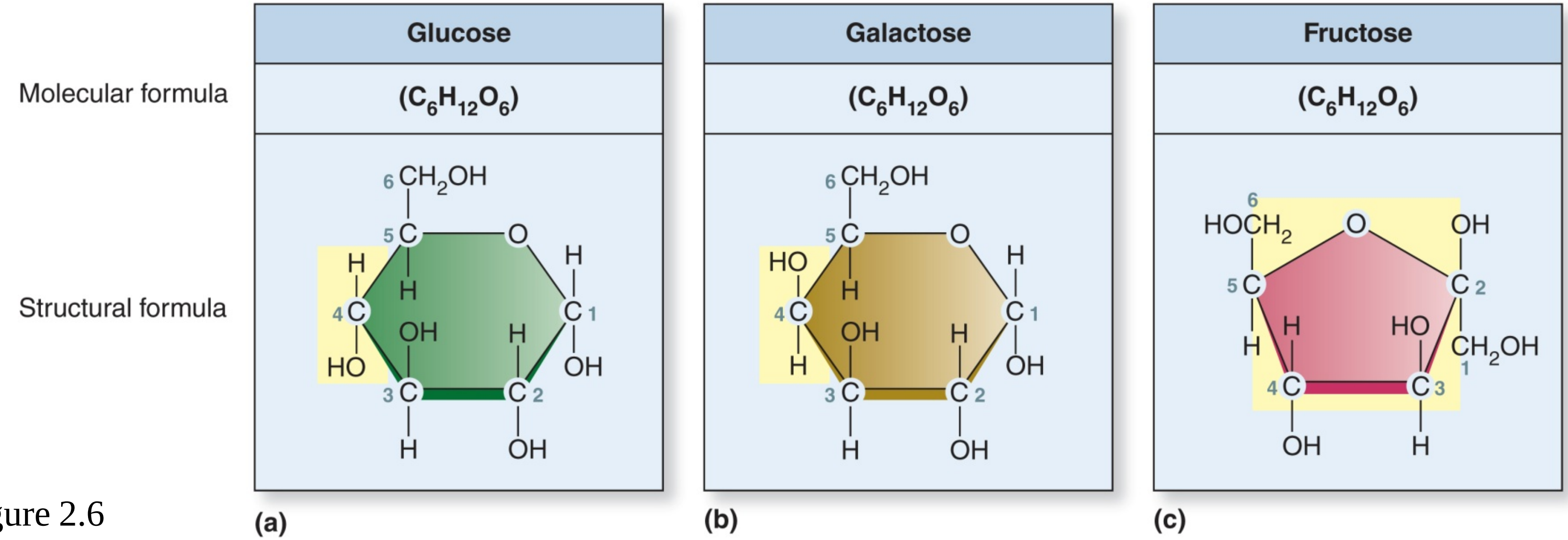


Figure 2.6

What did you learn?

- What information about a molecule is gained by a structural formula?
- How does this differ from a molecular formula?
- What is an isomer?

2.3b Covalent Bonds

- **Covalent bond**
  - Atoms share electrons
  - Occurs when both atoms require electrons
  - Occurs with atoms with 4 to 7 electrons in outer shell
  - Formed commonly in human body using
    - Hydrogen (H)
    - Oxygen (O)
    - Nitrogen (N)
    - Carbon (C)

2.3b Covalent Bonds

Number of covalent bonds an atom can form

- Simplest occurs between two hydrogen atoms
  - Each sharing its single electron
- Oxygen needs two electrons to complete outer shell
  - Forms two covalent bonds
- Nitrogen forms three bonds
- Carbon forms four bonds

2.3b Covalent Bonds

Single, double, and triple covalent bonds

- **Single covalent bond**
  - One pair of electrons shared
    - E.g., between two hydrogen atoms
- **Double covalent bond**
  - Two pairs of electrons shared
    - E.g., between two oxygen atoms
- **Triple covalent bond**
  - Three pairs of electrons shared
    - E.g., between two nitrogen atoms

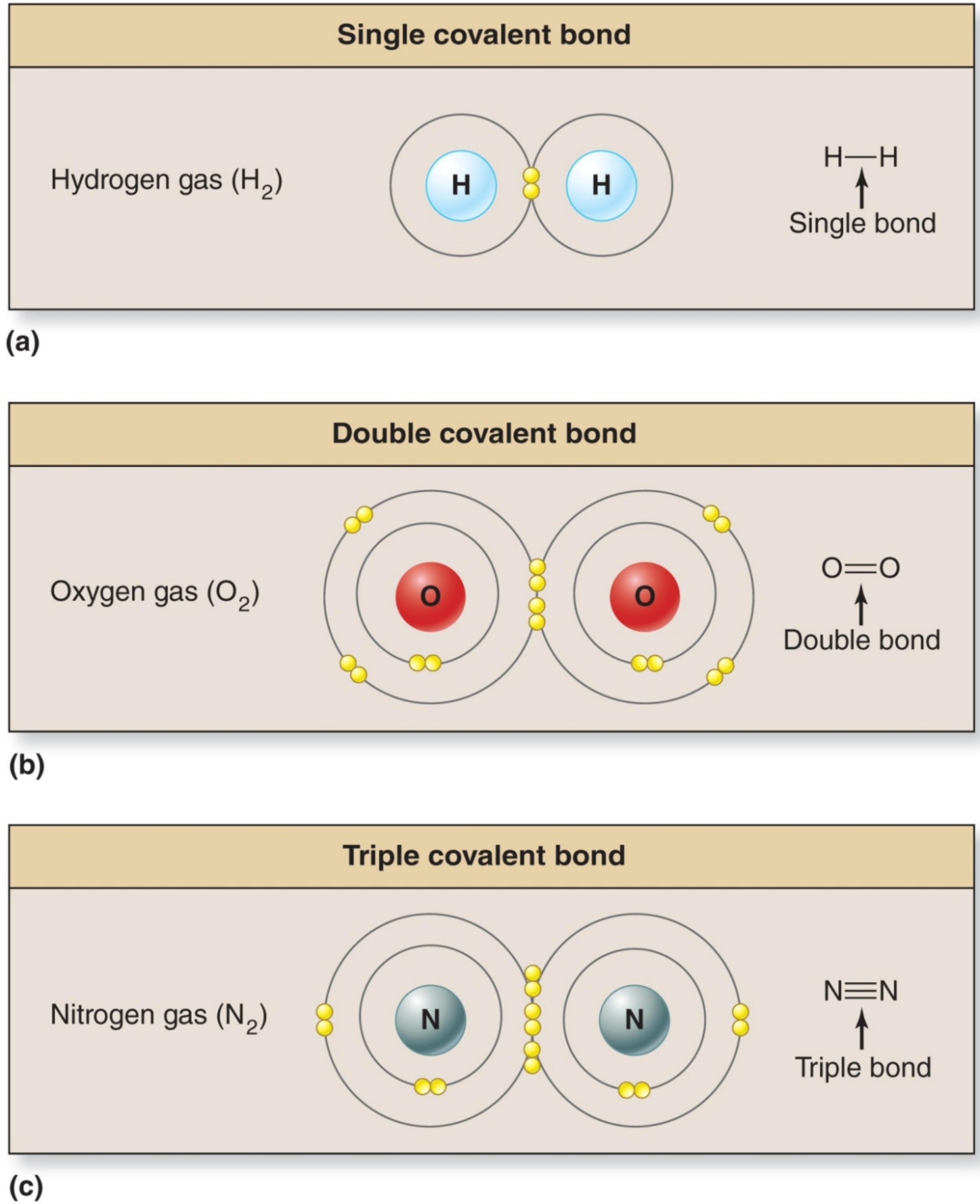

Single covalent bond
Hydrogen gas (H_2)
H
H
H—H
Single bond
(a)
Double covalent bond
Oxygen gas (O_2)
O
O
O═O
Double bond
(b)
Triple covalent bond
Nitrogen gas (N_2)
N
N
N≡N
Triple bond
(c)

Figure 2.7

2.3b Covalent Bonds

- Carbon needs four electrons to satisfy octet rule
  - Can be obtained in a number of ways

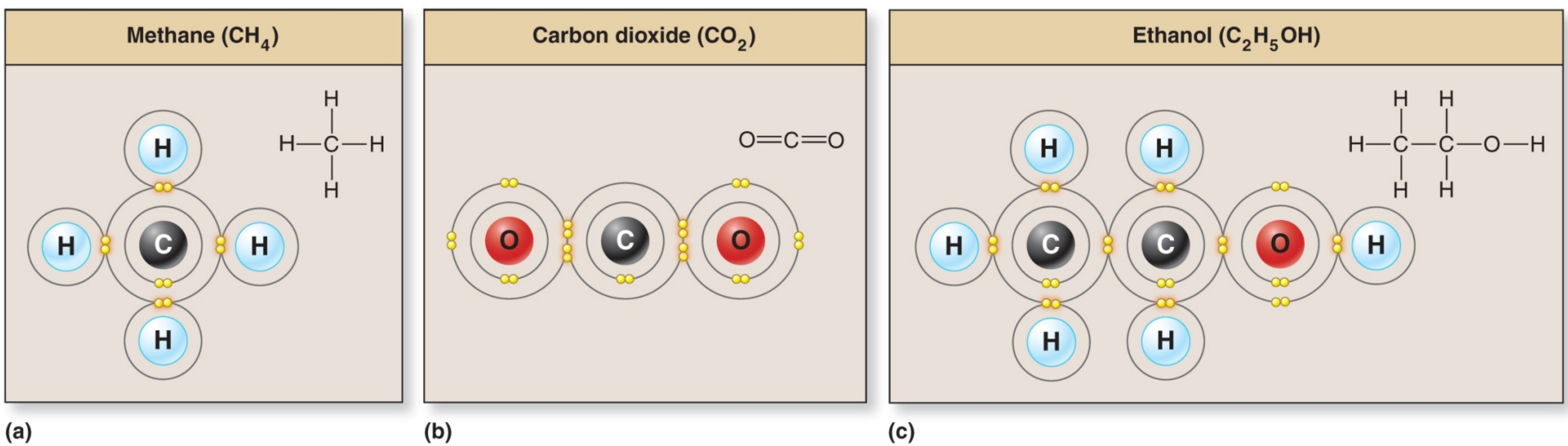


Figure 2.8

2.3b Covalent Bonds

Carbon skeleton formation

- Carbon
  - Bonds in straight chains, branched chains, or rings
  - Carbon present where lines meet at an angle
  - Additional atoms are hydrogen

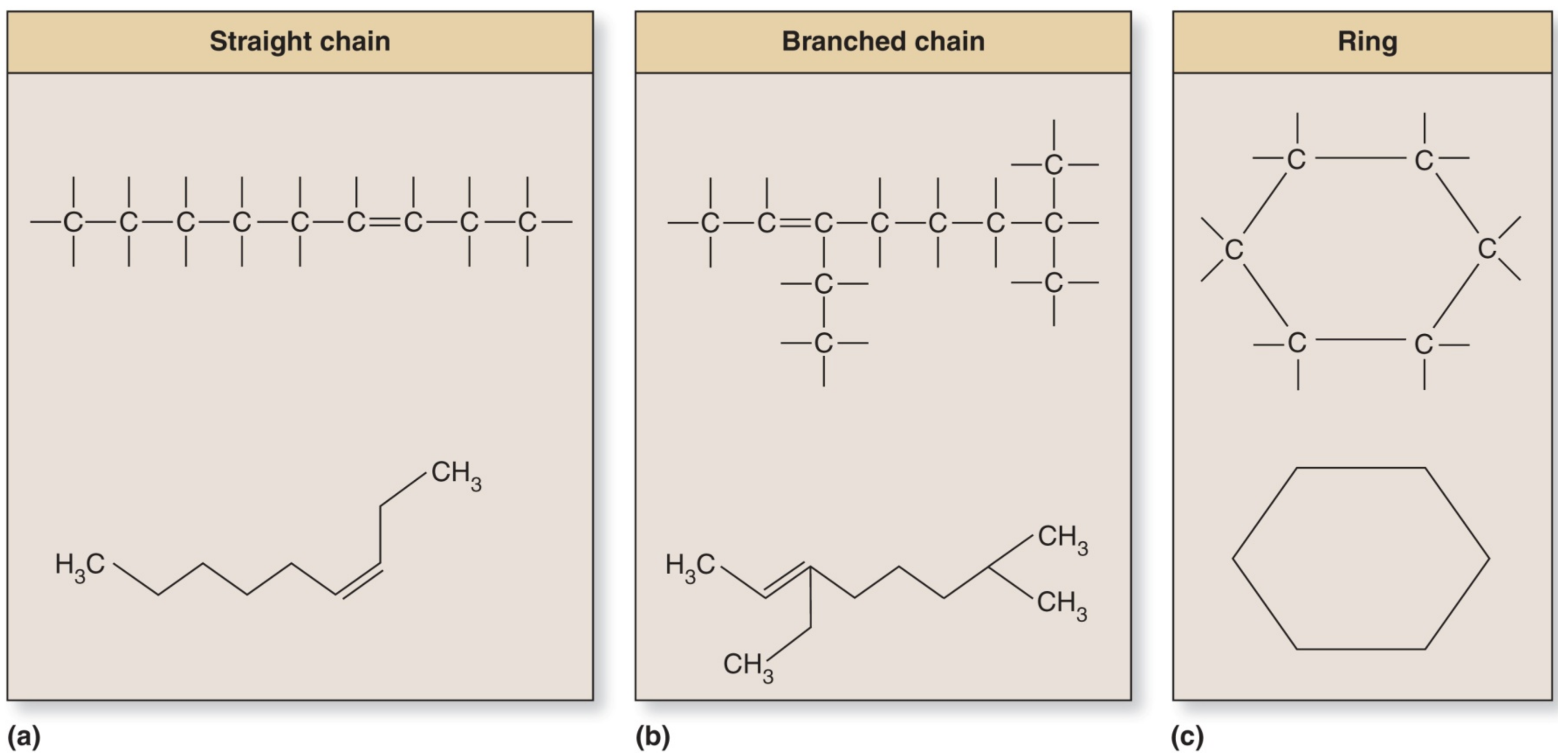


Figure 2.9

2.3b Covalent Bonds

Nonpolar and polar covalent bonds

- **Electronegativity—**relative attraction of each atom for electrons
  - Determines how electrons are shared in covalent bonds
  - Two atoms of same element have equal attraction for electrons
    - Resulting bond is **nonpolar covalent bond**
  - Sharing of electrons unequally = **polar covalent bond**

2.3b Covalent Bonds

Nonpolar and polar covalent bonds (*continued*)

- In periodic table, electronegativity increases
  - From left to right across row
  - From bottom to top in column
- For 4 most common elements composing living organisms
  - From least to greatest electronegativity
    - hydrogen < carbon < nitrogen < oxygen

2.3b Covalent Bonds

- Electrons have negative charge
  - More electronegative atom develops a partial negative charge
  - Less electronegative atom develops a partial positive charge
    - Written using Greek delta (δ) followed by superscript plus or minus
  - Exception to rule of polar bond forming between two different atoms
    - Carbon bonding with hydrogen

What did you learn?

- What type of bond is formed between two oxygen atoms?
- Why are some covalent bonds nonpolar and others polar?

2.3c Nonpolar, Polar, and Amphipathic Molecules

- Covalent bonds may be polar or nonpolar
  - **Nonpolar molecules** contain nonpolar covalent bonds
    - E.g., O—O and C—H are nonpolar bonds
  - **Polar** molecules contain polar covalent bonds
    - O—H is a polar bond in the polar molecule water (H2O)
  - Nonpolar molecules may contain polar covalent bonds, if the polar covalent bonds cancel each other
    - E.g., carbon dioxide

2.3c Nonpolar, Polar, and Amphipathic Molecules

- **Amphipathic molecules**
  - Large molecules with both polar and nonpolar regions
  - E.g., phospholipids

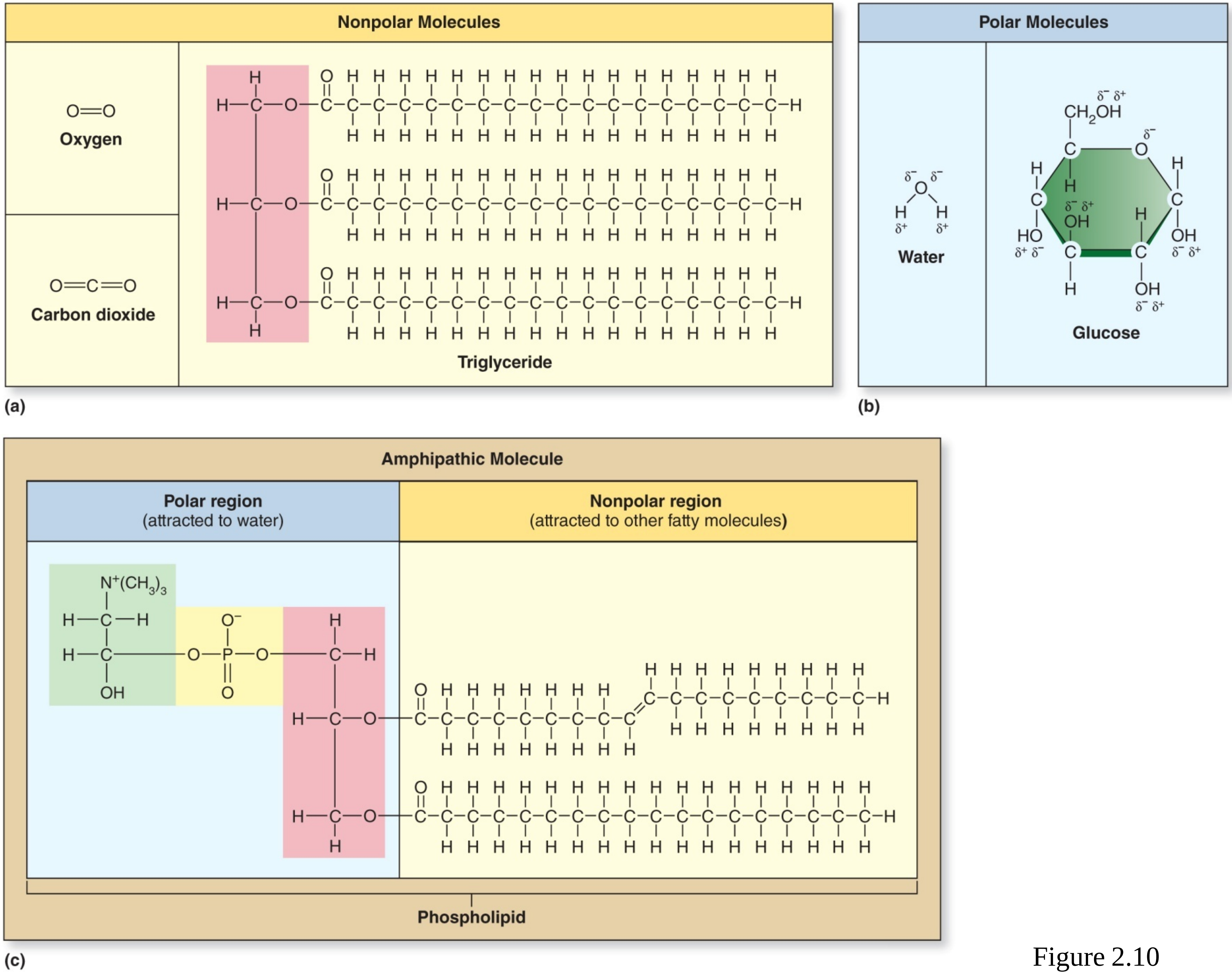

Nonpolar Molecules
Oxygen
Carbon dioxide
Triglyceride
(a)
Polar Molecules
Water
Glucose
(b)
Amphipathic Molecule
Polar region
(attracted to water)
Nonpolar region
(attracted to other fatty molecules)
Phospholipid
(c)

Figure 2.10

What did you learn?

- Is molecular oxygen nonpolar or polar?
- Is carbon dioxide polar or nonpolar?

2.3d Intermolecular Attractions

- **Intermolecular attractions**
  - Weak chemical attractions between molecules
  - Important for shape of complex molecules
    - E.g., DNA and proteins
  - **Hydrogen bond**
    - Forms between polar molecules
    - Attraction between partially positive hydrogen atom and a partially negative atom
    - Individually weak, collectively strong
    - Influences how water molecules behave

Hydrogen Bonding

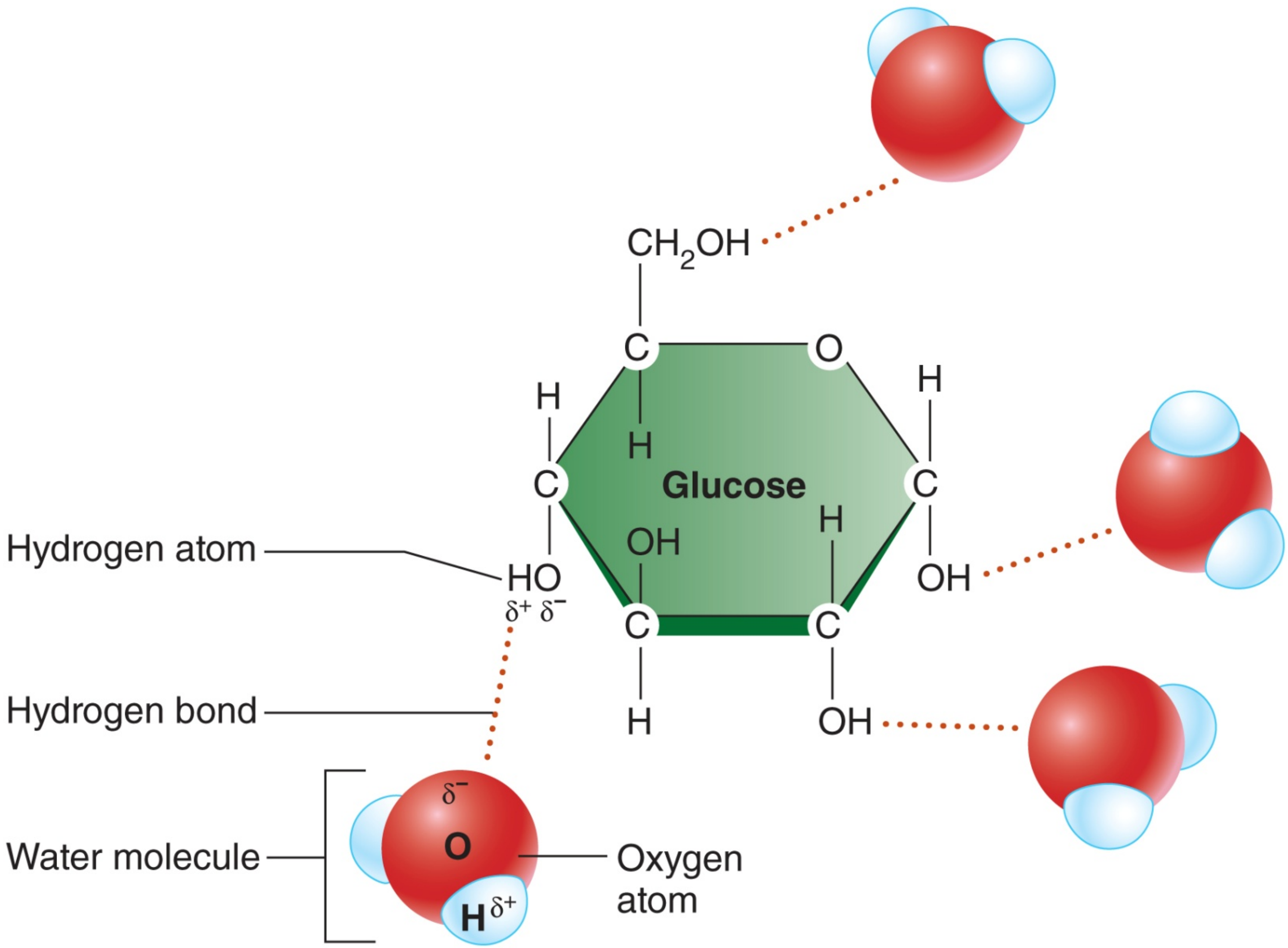


Figure 2.11

2.3d Intermolecular Attractions

- **Other intermolecular attractions**
  - **Van der Waals forces**
    - Nonpolar molecules
    - Electrons orbiting nucleus briefly, unevenly distributed
    - Induce unequal distribution of adjacent atom of another nonpolar molecule
    - Individually weak
  - **Hydrophobic interactions**
    - Nonpolar molecules placed in a polar substance
    - If occurring between parts of large molecule, termed **intramolecular attractions**

What did you learn?

- What is the name of the intermolecular attraction between a partially charged hydrogen of one polar molecule with a partially negative atom of another polar molecule?

2.4 Molecular Structure of Water and the Properties of Water

Learning Objectives:

1. Describe the molecular structure of water and how water molecules form four hydrogen bonds.
2. List the different properties of water and provide an example of the importance of each property within the body.
3. Compare substances that dissolve in water with those that both dissolve and dissociate in water. Distinguish between electrolytes and nonelectrolytes.
4. Describe the chemical interactions of nonpolar substances and water.
5. Explain how amphipathic molecules interact in water to form chemical barriers.

2.4a Molecular Structure of Water

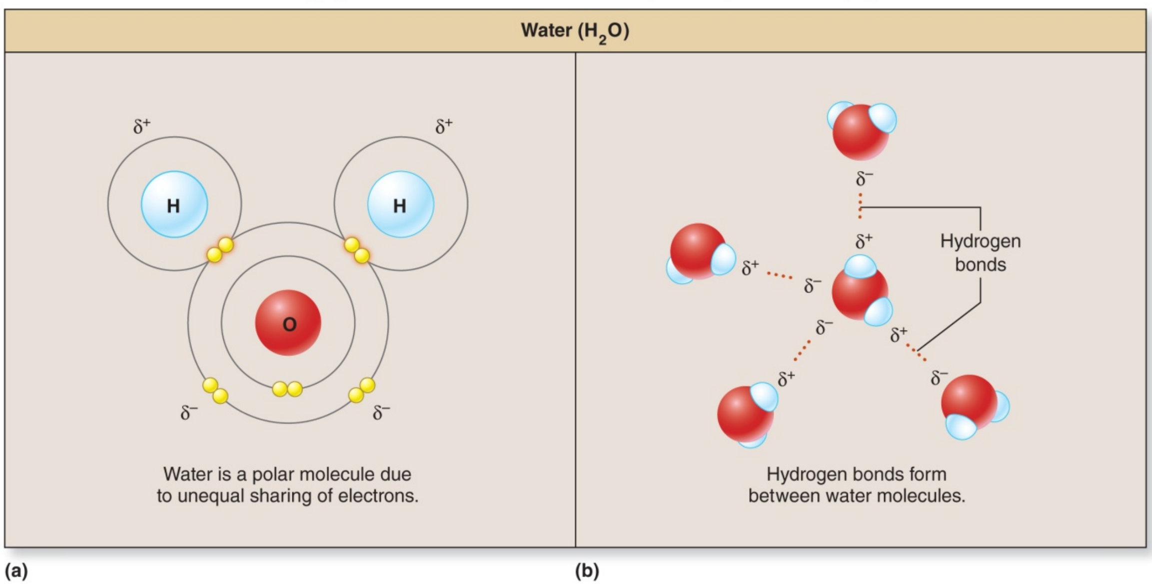


Figure 2.12

- Water
  - Composes two-thirds of the human body by weight
  - Polar molecule
    - One oxygen atom bonded to two hydrogen atoms
    - Oxygen atom has two partial negative charges
    - Hydrogens have single partial positive charge
  - Can form four hydrogen bonds with adjacent molecules
    - Central to water's properties

2.4b Properties of Water

Phases of water

- 3 phases of water, depending on temperature
  - Gas (water vapor)
    - Substances with low molecular mass
  - Liquid (water)
    - Almost all water in the body
    - Liquid at room temperature due to hydrogen bonding
  - Solid (ice)

2.4b Properties of Water

Phases of water (*continued*)

- Functions of liquid water
  - Transports
    - Substances dissolved in water move easily throughout body
  - Lubricates
    - Decreases friction between body structures
  - Cushions
    - Absorbs sudden force of body movements
  - Excretes wastes
    - Unwanted substances dissolve in water are easily eliminated

2.4b Properties of Water

- **Cohesion**
  - Attraction between water molecules due to hydrogen bonding
- **Surface tension**
  - Inward pulling of cohesive forces at surface of water
  - Causes moist sacs of air in lungs to collapse
    - Surfactant, a lipoprotein, prevents collapse
- **Adhesion**
  - Attraction between water molecules and a substance other than water

2.4b Properties of Water

High specific heat and high heat of vaporization

- **Temperature**
  - Measure of kinetic energy of atoms or molecules within a substance
- **Specific heat**
  - Amount of energy required to increase temperature of 1 gram of a substance by 1 degree Celsius
  - Water’s value extremely high due to energy needed to break hydrogen bonds
  - Contributes to keeping body temperature constant

2.4b Properties of Water

High specific heat and high heat of vaporization (*continued*)

- **Heat of vaporization**
  - Heat required for release of molecules from a liquid phase into a gaseous phase for 1 gram of a substance
  - Water's value very high due to hydrogen bonding
  - Sweating cools body
    - Excess heat dissipated as water evaporates

What did you learn?

- Which property of water contributes to the need to produce surfactant to prevent the collapse of the alveoli?
- Which property contributes to body temperature regulation through sweating?
- Why is sweating less effective in cooling the body on a humid day?

2.4c Water as the Universal Solvent

- Water—**solvent** of the body
- **Solutes** are substances that dissolve in water
- Water is called **universal solvent** because most substances dissolve in it
  - Chemical properties of a substance determine whether it will dissolve or not

2.4c Water as the Universal Solvent

Substances that dissolve in water

- Polar molecules and ions
  - **Hydrophilic** means “water-loving”
  - Water surrounds substances, forms a **hydration shell**
  - Some substances dissolve but remain intact
    - E.g., glucose and alcohol
    - **Nonelectrolytes** remain intact but do not conduct current
  - Substances dissolve and **dissociate** (separate)
    - NaCl dissociates into Na^+ and Cl^- ions
    - Acids and bases, such as HCl
    - **Electrolytes** can conduct current

2.4c Water as the Universal Solvent

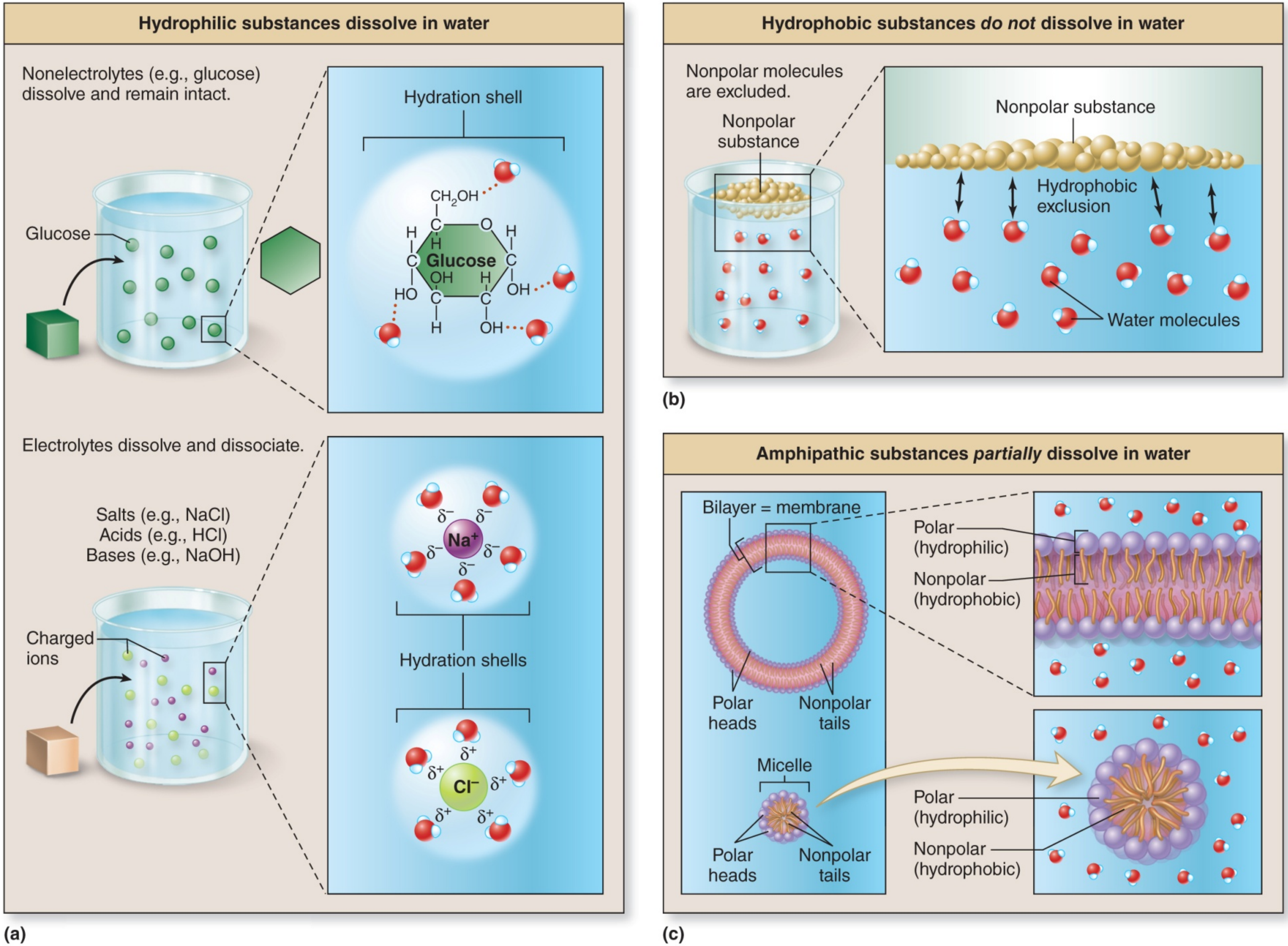
Substances that do not dissolve in water

- Nonpolar molecules
  - **Hydrophobic** means “water-fearing”
  - **Hydrophobic exclusion—**cohesive water molecules “force out” nonpolar molecules
  - **Hydrophobic interaction—**“excluded molecules”
  - Hydrophobic substances require carrier proteins to be transported within the blood
    - E.g., fats and cholesterol are unable to dissolve within water

2.4c Water as the Universal Solvent

Substances that partially dissolve in water

- Amphipathic molecules have polar and nonpolar regions
  - Polar portion of molecule dissolves in water
  - Nonpolar portion repelled by water
- Phospholipid molecules are amphipathic
  - Polar heads have contact with water
  - Nonpolar tails group together
  - Results in bilayers of phospholipid molecules
    - E.g., membranes of a cell
- Other amphipathic molecules form a *micelle*


Hydrophilic substances dissolve in water
Nonelectrolytes (e.g., glucose) dissolve and remain intact.
Glucose
Hydration shell
CH_2OH
Glucose
Electrolytes dissolve and dissociate.
Salts (e.g., NaCl)
Acids (e.g., HCl)
Bases (e.g., NaOH)
Charged ions
Na^+
δ^-
Hydration shells
Cl^-
δ^+
(a)
Hydrophobic substances *do not* dissolve in water
Nonpolar molecules are excluded.
Nonpolar substance
Nonpolar substance
Hydrophobic exclusion
Water molecules
(b)
Amphipathic substances *partially* dissolve in water
Bilayer = membrane
Polar heads
Nonpolar tails
Micelle
Polar heads
Nonpolar tails
Polar (hydrophilic)
Nonpolar (hydrophobic)
Polar (hydrophilic)
Nonpolar (hydrophobic)
(c)

2.5 Acidic and Basic Solutions, pH, and Buffers

Learning Objectives:

1. Describe what is formed when water dissociates.
2. Explain the difference between an acid and a base.
3. Define pH and explain the relative pH values of both acids and bases.
4. Explain the term neutralization, and describe how the neutralization of both an acid and a base occur.
5. Describe the action of a buffer.

2.5a Water: A Neutral Solvent

- Water spontaneously dissociates to form ions
  - Bond between oxygen and hydrogen breaks apart spontaneously
    - 1/10,000,000 ions per liter
    - OH group hydroxide ion (OH^-)
  - Hydrogen ion transferred to a second water molecule
    - Hydronium ion (H_3O+)
  - Equal numbers of positive hydrogen ions and negative hydroxyl ions produced
    - Water remains neutral

$$H_2O + H_2O \longrightarrow H_3O^+ + OH^-$$

simplified to

$$H_2O \longrightarrow H^+ + OH^-$$

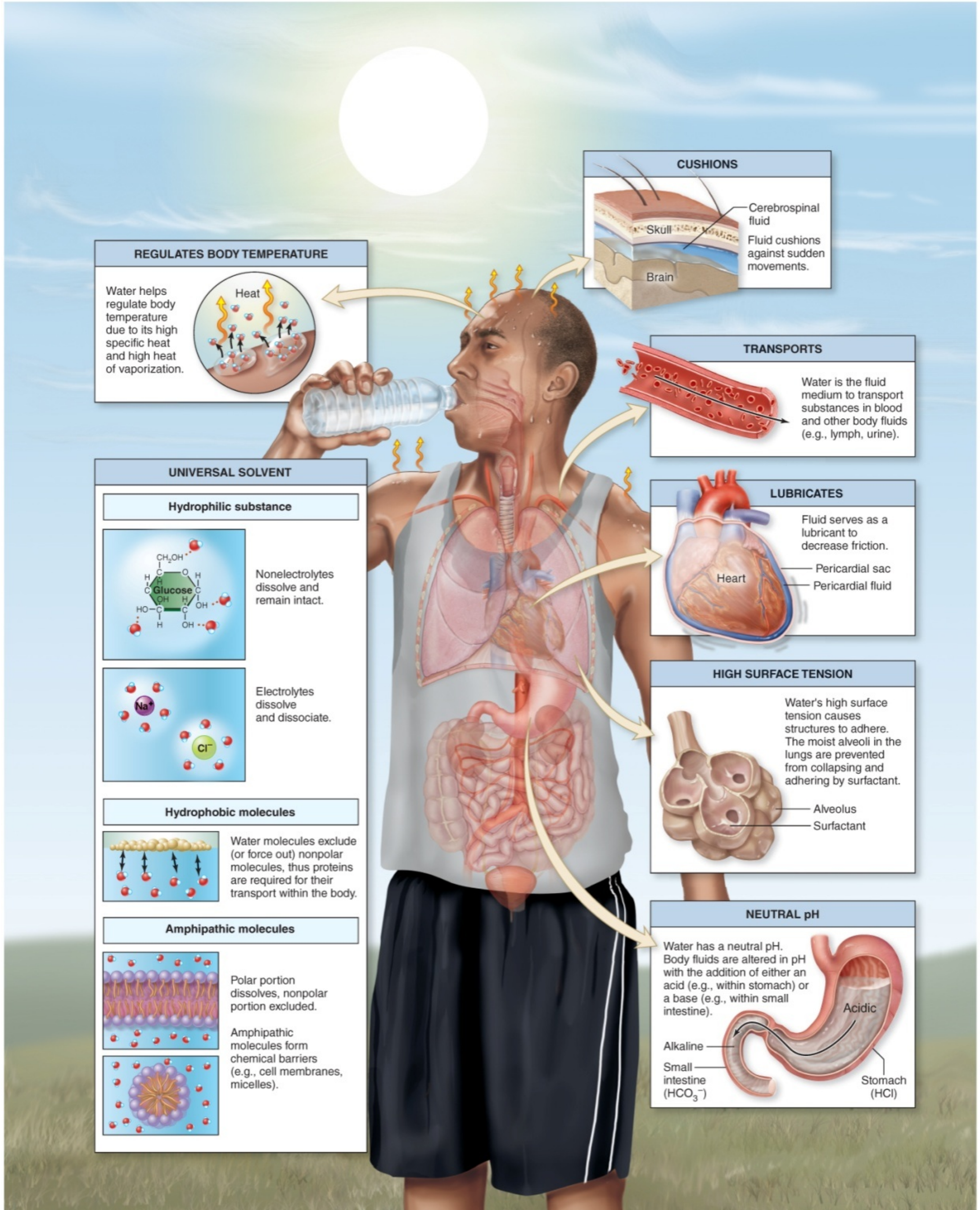

CUSHIONS
Cerebrospinal fluid
Skull
Brain
Fluid cushions against sudden movements.
REGULATES BODY TEMPERATURE
Water helps regulate body temperature due to its high specific heat and high heat of vaporization.
Heat
TRANSPORTS
Water is the fluid medium to transport substances in blood and other body fluids (e.g., lymph, urine).
UNIVERSAL SOLVENT
Hydrophilic substance
Glucose
Nonelectrolytes dissolve and remain intact.
Electrolytes dissolve and dissociate.
Hydrophobic molecules
Water molecules exclude (or force out) nonpolar molecules, thus proteins are required for their transport within the body.
Amphipathic molecules
Polar portion dissolves, nonpolar portion excluded.
Amphipathic molecules form chemical barriers (e.g., cell membranes, micelles).
LUBRICATES
Fluid serves as a lubricant to decrease friction.
Heart
Pericardial sac
Pericardial fluid
HIGH SURFACE TENSION
Water's high surface tension causes structures to adhere. The moist alveoli in the lungs are prevented from collapsing and adhering by surfactant.
Alveolus
Surfactant
NEUTRAL pH
Water has a neutral pH. Body fluids are altered in pH with the addition of either an acid (e.g., within stomach) or a base (e.g., within small intestine).
Acidic
Alkaline
Small intestine (HCO_3^-)
Stomach (HCl)

Figure 2.14

What did you learn?

- How does the interaction of a nonelectrolyte and water differ from the interaction of an electrolyte and water?

- Why is water neutral?

2.5b Acids and Bases

- **Acid** dissociates in water to produce H^+ and an anion
  - **Proton donor**
  - Increases concentration of free H^+
  - More dissociation of H^+ with stronger acids
    - E.g., HCl in the stomach
  - Less dissociation of H^+ with weaker acids
    - E.g., carbonic acid in the blood

$$\text{Substance A (an acid in water)} \longrightarrow H^+ + \text{Anion}$$

2.5b Acids and Bases

- **Base** accepts H^+ when added to solution
  - **Proton acceptor**
  - Decreases concentration of free H^+
  - More absorption of H^+ with stronger bases
    - E.g., ammonia and bleach
  - Less absorption of H^+ with weaker bases
    - E.g., bicarbonate in blood and in secretions released into small intestine

Substance B (a base in water) + $H^+ \longrightarrow$ B—H

2.5c pH, Neutralization, and the Action of Buffers

- pH is a measure of H^+
  - Relative amount of H^+ in a solution
  - Range between 0 and 14
- The pH of plain water is 7
  - Water dissociates to produce 1/10,000,000 of H^+ and OH^- ions per liter
  - Equal to 1×10^{-7} or to 0.0000001
- pH and H^+ concentration are inversely related
  - Inverse of the log for a given H^+ concentration
  - As H^+ concentration increases, pH decreases
  - As H^+ concentration decreases, pH increases

2.5c pH, Neutralization, and the Action of Buffers

Interpreting the pH scale

- Solutions with equal concentrations of H^+ and OH^-
  - Are neutral
  - Have a pH of 7
- Solutions with greater H^+ than OH^-
  - Are acidic
  - Have a pH < 7
- Solutions with greater OH^- than H^+
  - Are basic (alkaline)
  - Have a pH > 7
- Moving from one increment to next is a 10-fold change
  - E.g., a pH of 6 has 10 times greater concentration of H^+ than pure water

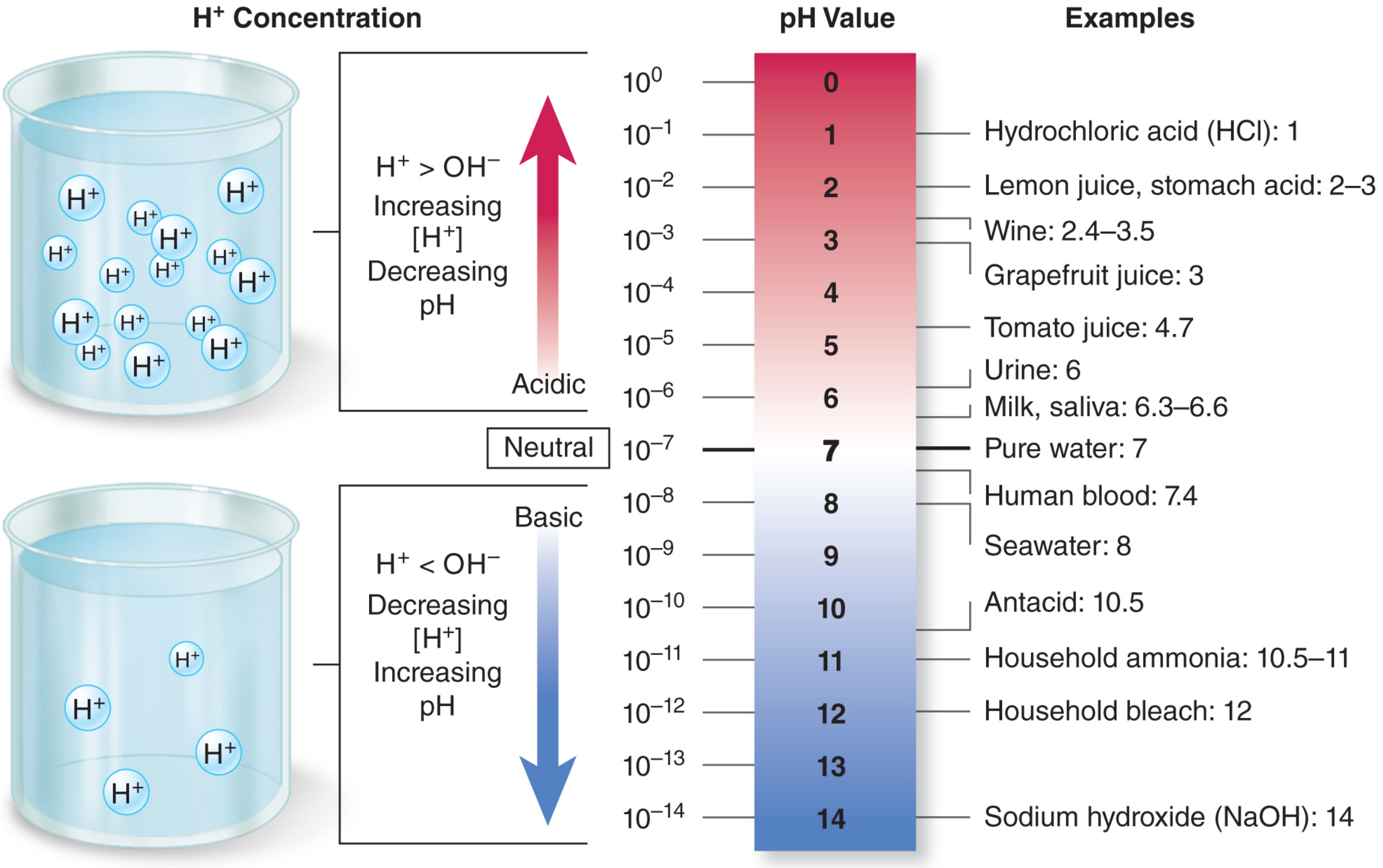

H+ Concentration
pH Value
Examples
H+ > OH−
Increasing [H+]
Decreasing pH
Acidic
Neutral
Basic
H+ < OH−
Decreasing [H+]
Increasing pH
10^0 0
10^−1 1
10^−2 2
10^−3 3
10^−4 4
10^−5 5
10^−6 6
10^−7 7
10^−8 8
10^−9 9
10^−10 10
10^−11 11
10^−12 12
10^−13 13
10^−14 14
Hydrochloric acid (HCl): 1
Lemon juice, stomach acid: 2–3
Wine: 2.4–3.5
Grapefruit juice: 3
Tomato juice: 4.7
Urine: 6
Milk, saliva: 6.3–6.6
Pure water: 7
Human blood: 7.4
Seawater: 8
Antacid: 10.5
Household ammonia: 10.5–11
Household bleach: 12
Sodium hydroxide (NaOH): 14

Figure 2.15

2.5c pH, Neutralization, and the Action of Buffers

- **Neutralization**
  - When an acidic or basic solution is returned to neutral (pH 7)
  - Acids neutralized by adding base
    - E.g., medications to neutralize stomach acid must contain a base
  - Bases neutralized by adding acid
- **Buffers**
  - Help prevent pH changes if excess acid or base is added
  - Act to accept H^+ from excess acid or donate H^+ to neutralize base
    - Carbonic acid (weak acid) and bicarbonate (weak base) buffer blood pH
    - Both help maintain blood pH in a critical range (7.35 to 7.45)

What did you learn?

- What type of substances release H^+ when added to water?
- What is the general relationship of H^+ and pH?
- Why are buffers important?

2.6
Water Mixtures

Learning Objectives:

1. Compare and contrast the three different types of water mixtures.
2. Explain how an emulsion differs from other types of mixtures.
3. Explain the different ways to express the concentration of solute in a solution.

2.6 Water Mixtures

- **Mixtures** are formed from combining two or more substances
- Two defining features
  - Substances mixed are not chemically changed
  - Substances can be separated by physical means
    - E.g., evaporation or filtering

2.6a Categories of Water Mixtures

- Three categories of water mixtures
  - **Suspension:** material larger in size than 100 nanometers mixed with water
    - E.g., blood cells within plasma or sand in water
    - Does not remain mixed unless in motion
    - Appears cloudy or opaque; scatters light
  - **Colloid:** protein of size from 1to100 nanometers
    - E.g., fluid in cell cytosol and fluid in blood plasma
    - Remains mixed when not in motion
    - Scatters light

2.6a Categories of Water Mixtures

- Three categories of water mixtures (*continued*)
  - **Solution:** homogeneous mixture of material smaller than 1 nanometer
    - Dissolves in water
    - Does not scatter light; does not settle if solution not in motion
    - E.g., sugar water, salt water, blood plasma
  - Special category of suspension—**emulsion**
    - Water and a nonpolar liquid substance
    - E.g., oil and vinegar salad dressing or breast milk
    - Does not mix unless shaken

Mixtures and Emulsions

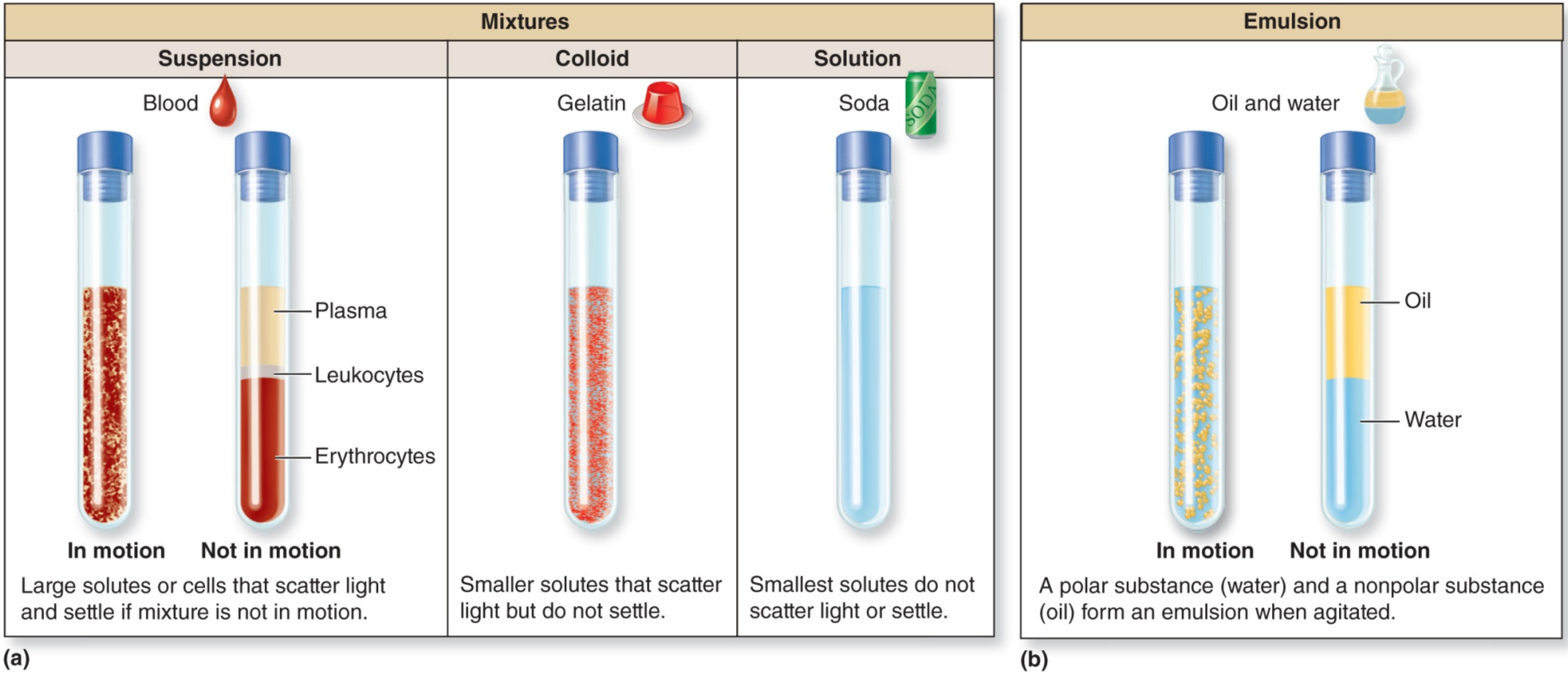


Figure 2.16

2.6b Expressions of Solution Concentration

- Concentration is determined by the amount of solute dissolved in a solution
- Expressions of concentration
  - **Mass/volume**
    - Mass of solute per volume of solution
    - E.g., results from a blood test
  - **Mass/volume percent**
    - Grams of solute per 100 mL solution
    - E.g., IV solutions

2.6b Expressions of Solution Concentration

- Expressions of concentration (*continued*)
  - **Molarity**
    - Moles solute/L solution
    - Alters with changes in temperature
    - More easily measured in the body than molality
  - **Molality**
    - Moles solute/kg solvent
    - Does not alter with changes in temperature
    - Slightly more accurate than molarity

Table 2.2 Expressing Solution Concentrations

Solution Concentration	Expressed As	Unit of Measurement	Examples
Mass/volume	Mass of solute per volume of solution	μg solute/dL solution mg solute/dL solution	Normal blood concentration of iron is within the range of 40 to 150 μg/dL. Normal blood concentration of glucose is between 70 and 110 mg/dL.
Mass/volume percent	Grams of solute per 100 milliliters (mL) of solution	grams/100 mL	5% dextrose intravenous (IV) solution (D5W) has a concentration of 5 grams of dextrose (glucose) per 100 mL of solution. Physiologic saline (0.9% NaCl) has 0.9 grams of NaCl per 100 mL of solution.
Molarity	Moles of solute per liter of solution	moles solute/L solution	0.164 mol/L solution
Molality	Moles of solute per kilogram of solvent	moles solute/kg solvent	0.164 mol/kg solvent

mg = milligrams; dL = deciliters; kg = kilograms; μg = micrograms

2.6b Expressions of Solution Concentration

- **Osmoles (osm)** is the unit of measurement for the number of particles in a solution
  - Reflects whether a substance dissolves, or dissolves and dissociates
  - Can be expressed as osmolarity or osmolality
  - **Osmolarity**
    - Number of particles in a 1 liter solution
    - Easier to measure
    - Blood serum expresses as milliosmoles
  - **Osmolality**
    - Number of particles in 1 kg of water
    - More accurate, but difficult to measure

2.6b Expressions of Solution Concentration

- **Mole** = 6.022×10^{23} atoms, ions, or molecules
  - Mass in grams equal to atomic mass of an element or molecular mass of a compound
    - E.g., one mole carbon = 12.01 grams
- To find **molecular mass,** multiply number of units of each element by average atomic mass and add totals
  - Some variation due to isotopes

What did you learn?

- A water mixture of unknown type shows settling of materials to the bottom of the tube when it is not in motion. What type is it?

- What are the four ways solution concentration may be expressed?

2.7 Biological Macromolecules

Learning Objectives:

1. Differentiate between an organic molecule and an inorganic molecule.
2. Describe the general chemical composition of biomolecules.
3. Define a monomer and polymer.
4. Describe the role of water in both dehydration and hydrolysis reactions in altering biomolecules.
5. Describe the general characteristics of a lipid.

2.7 Biological Macromolecules (*continued*)

Learning Objectives:

6. Identify the four types of lipids and their physiologic roles.
7. Describe the distinguishing characteristics of carbohydrates.
8. Explain the relationship between glucose and glycogen.
9. Name some other carbohydrates found in living systems.
10. Describe the general structure of a nucleic acid.

2.7 Biological Macromolecules (*continued*)

Learning Objectives:

11. Describe the structure of a nucleotide monomer.
12. Distinguish between DNA and RNA.
13. Name other important nucleotides.
14. List the general functions of proteins.
15. Describe the general structure of amino acids and proteins.

2.7a Biological Macromolecules: General Characteristics

- **Organic molecules—**molecules that contain carbon
  - Most are a component of living organisms
  - Biological macromolecules (*biomolecules*) are a subset
- **Inorganic molecules—**all other molecules
- Four classes of organic biomolecules in living systems
  1. Lipids
  2. Carbohydrates
  3. Nucleic acids
  4. Proteins

2.7a Biological Macromolecules: General Characteristics

- Chemical composition
  - Biomolecules always contain carbon, hydrogen, and generally oxygen
  - Some may also have nitrogen, phosphorus, or sulfur
  - Clustered on right side of periodic table
- "Carbon skeletons" can take a variety of forms
- Hydrocarbons
  - Contain only carbon and hydrogen atoms
  - Nonpolar and hydrophobic
  - E.g., methane gas

2.7a Biological Macromolecules: General Characteristics

- Chemical composition (*continued*)
  - Carbon skeleton may contain two or more atoms with specific characteristics
  - Functional groups
  - Most are polar and able to hydrogen bond
  - Some act like acids (e.g., carboxyl group)
  - Others act like bases (e.g., amine group)

2.7a Biological Macromolecules: General Characteristics

- **Polymers**
  - Molecules made of **monomers** (repeating subunits)
    - Monomers are identical or similar in chemical structure
    - E.g., carbohydrates, nucleic acids, proteins
    - Carbohydrates contain sugar monomers
    - Nucleic acids contain nucleotide monomers
    - Proteins contain amino acid monomers

2.7a Biological Macromolecules: General Characteristics

- **Dehydration synthesis** (*condensation*)
  - Occurs during the synthesis of biomolecules
  - One subunit looses an —H
  - Other subunit loses an —OH
  - New covalent bond formed and water produced
- **Hydrolysis**
  - Occurs during the breakdown of biomolecules
  - An —H added to one subunit
  - An —OH added to another subunit

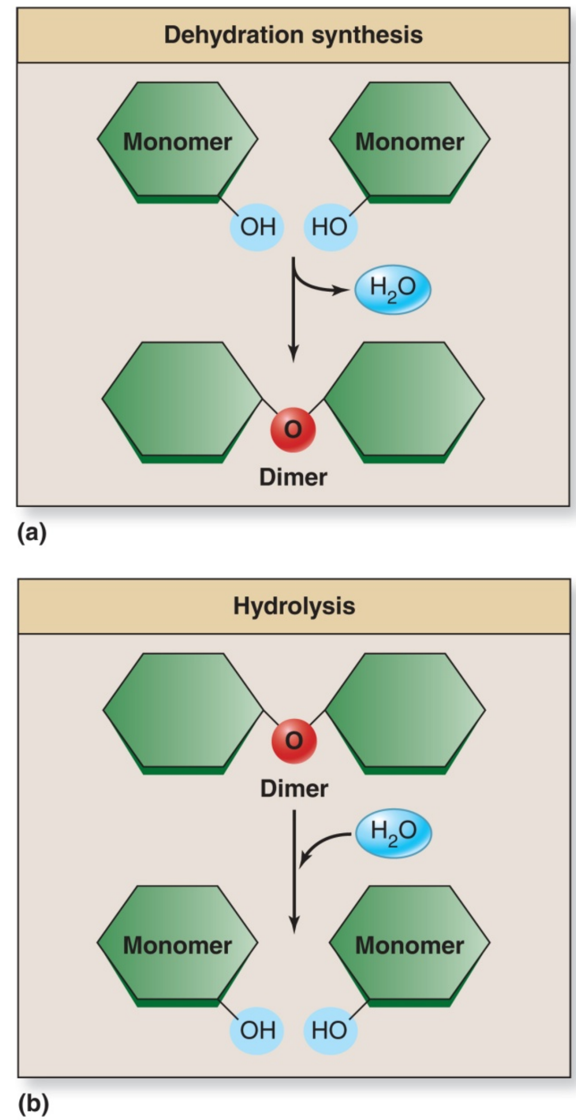


Figure 2.17

What did you learn?

- What are the three biomolecules that are polymers?
- What are the monomers that compose them?

2.7b Lipids

- Lipids
  - Only category of biomolecules that are not polymers
  - Diverse group of fatty, water-insoluble molecules
  - Function as stored energy, cellular membrane components, hormones
  - Four primary classes
    1. Triglycerides
    2. Phospholipids
    3. Steroids
    4. Eicosanoids

2.7b Lipids

- **Triglycerides**
  - Most common form of lipid in living things
  - Long-term energy storage in adipose tissue
  - Structural support, cushioning, insulation
  - Formed from a glycerol molecule and three fatty acids
    - —H from the glycerol, an —OH from the fatty acid
  - Formed during a dehydration synthesis

2.7b Lipids

- Fatty acids
  - Vary in length
  - Vary in number of double bonds
    - **Saturated,** lack double bonds
    - **Unsaturated,** one double bond
    - **Polyunsaturated,** two or more double bonds
- Adipose tissue
  - **Lipogenesis—**formation of triglycerides when conditions of excess nutrients exist
  - **Lipolysis—**breakdown of triglycerides when nutrients are needed

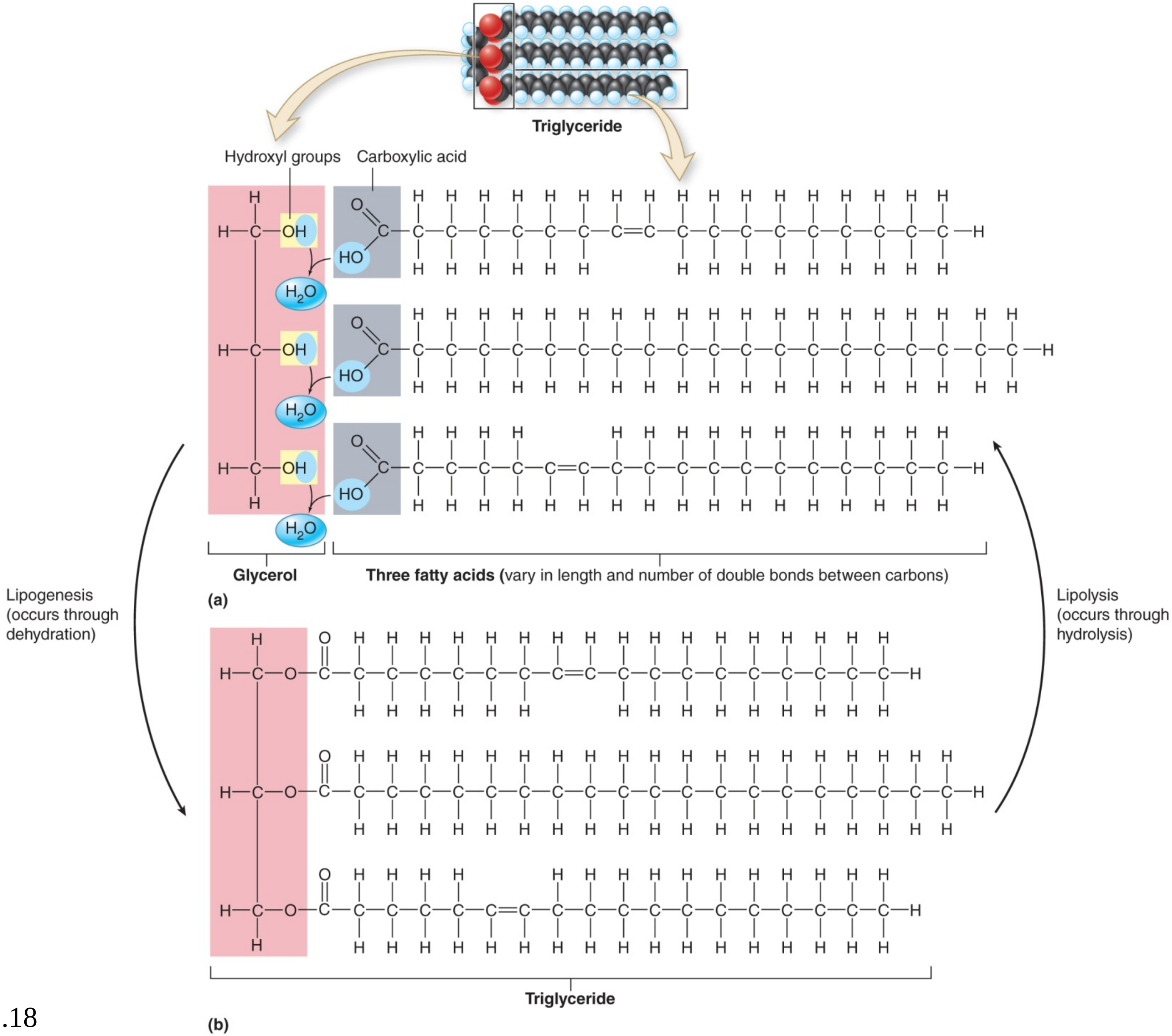

Triglyceride
Hydroxyl groups
Carboxylic acid
Glycerol
Three fatty acids (vary in length and number of double bonds between carbons)
(a)
Lipogenesis (occurs through dehydration)
Lipolysis (occurs through hydrolysis)
Triglyceride
(b)

Figure 2.18

2.7b Lipids

- **Phospholipids**
  - Amphipathic molecules that form chemical barriers of cell membranes
  - Phospholipid structure similar to a triglyceride
    - Except one end of the glycerol has a polar phosphate group with various organic groups attached instead of a fatty acid
    - Glycerol, phosphate, and organic groups are polar
      - Form **hydrophilic head**
    - Fatty acid group is nonpolar
      - Form **hydrophobic tails**

2.7b Lipids

- **Steroids**
  - Composed of hydrocarbons arranged in multiringed structure
    - Four carbon rings; three have 6 carbon atoms, one has 5 carbon atoms
  - Differ in side chains extending from their rings
    - Cholesterol
      - Component of animal plasma membranes
      - Precursor to other steroid synthesis
    - Steroid hormones (e.g., testosterone and estrogen)
    - Bile salts

2.7b Lipids

- **Eicosanoids**
  - Modified 20-carbon fatty acids
  - Synthesized from arachidonic acid, membrane component
  - Local signaling molecules
  - Primary functions in inflammatory response and nervous system communication
  - Four classes
    1. Prostaglandins
    2. Prostacyclins
    3. Thromboxanes
    4. Leukotrienes

2.7b Lipids

- Other lipids
  - Glycolipids
    - Lipid molecules with carbohydrate attached
    - Associated with plasma membrane
    - Involved in cellular binding to form tissues
  - Fat soluble vitamins
    - Vitamins A, E, and K

What did you learn?

- Which class of lipids forms cell membranes?

Clinical View—Fatty Acids: Saturated, Unsaturated, and Trans Fats

- Most animal fats are **saturated**
  - Most are solid at room temperature
- Most vegetable fats are **unsaturated**
  - Most are liquid at room temperature
  - Generally healthier
  - Can be converted to saturated fats through **hydrogenation**
- Partial hydrogenation may lead to **trans fats**
  - Increase the risk of heart attack and stroke

2.7c Carbohydrates

- **Carbohydrates**
  - An —H and an — OH are usually attached to every carbon
  - Chemical formula is $(CH_2O)_n$
    - o n = number of carbon atoms
  - **Monosaccharides**
    - o Simple sugar monomers
  - **Disaccharides**
    - o Formed from two monosaccharides
  - **Polysaccharides**
    - o Formed from many monosaccharides

2.7c Carbohydrates

- **Glucose**
  - Six-carbon carbohydrate
  - Most common monosaccharide
  - Primary nutrient supplying energy to cells
  - Concentration must be carefully maintained
- **Glycogen**
  - Liver and skeletal muscle store excess glucose, then bind glucose monomers together (**glycogenesis**)
  - Liver hydrolyzes glycogen into glucose as needed (**glycogenolysis**)
  - Liver can also form glucose from noncarb sources (**gluconeogenesis**)

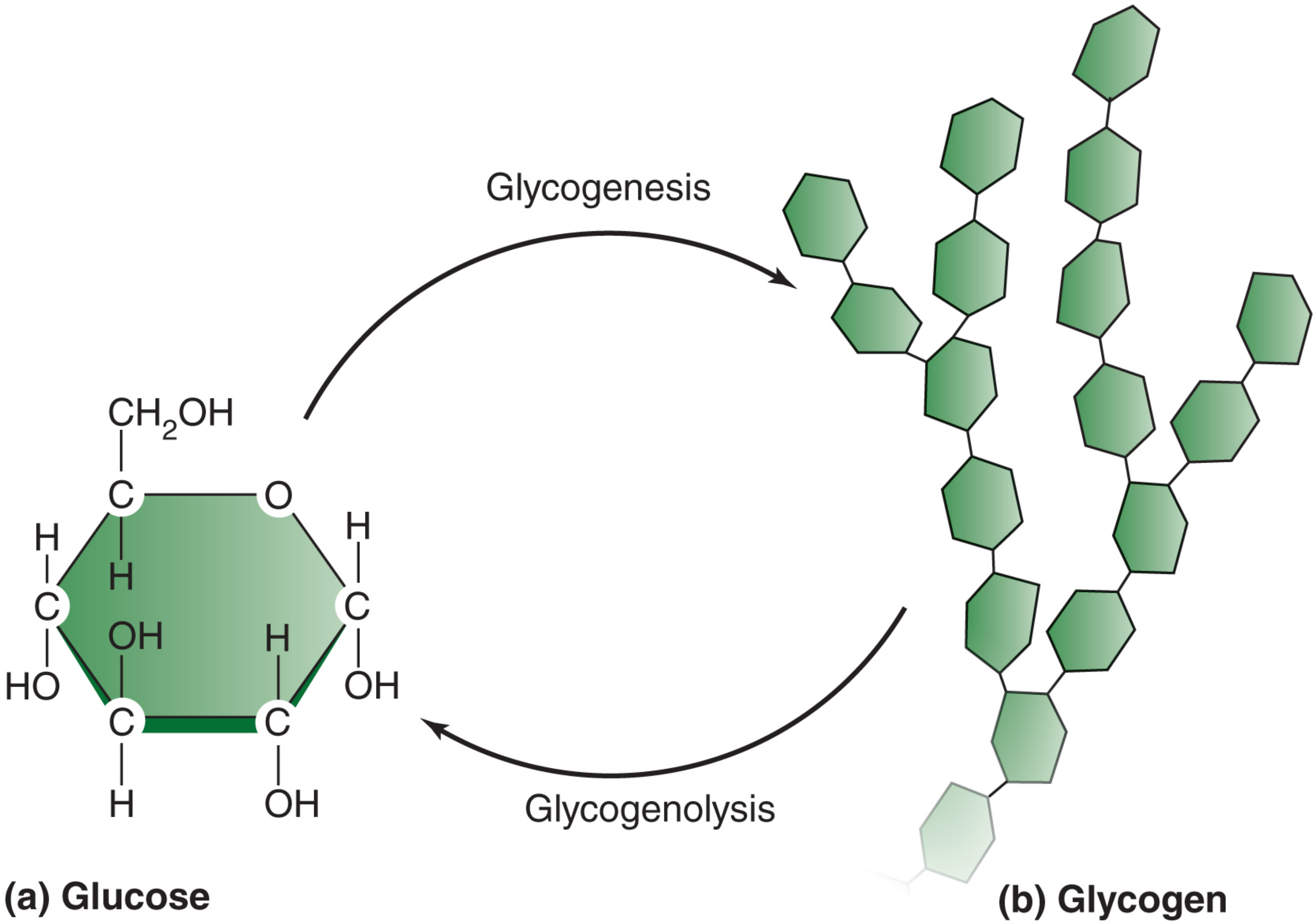

Glycogenesis
CH2OH
C
O
H
H
H
C
C
OH
H
HO
OH
C
C
H
OH
Glycogenolysis
(a) Glucose
(b) Glycogen

Figure 2.19

2.7c Carbohydrates

- Other types of carbohydrates
  - Hexose monosaccharides
    - Glucose isomers (e.g., galactose and fructose)
  - Five-carbon monosaccharides (pentose sugars)
    - E.g., ribose and deoxyribose
  - Disaccharides—two simple sugars bonded together
    - Most common are sucrose (table sugar), lactose (milk sugar), and maltose (malt sugar)
      - All with glucose bonded to a second hexose monosaccharide

2.7c Carbohydrates

- Other types of carbohydrates (*continued*)
  - Polysaccharides—three or more sugars
    - Glycogen most common in animals
    - Starch and cellulose found in plants
    - Plant starch is a major nutritional source of glucose for humans
    - Cellulose is a source of fiber (nondigestible)

Other Simple Carbohydrates

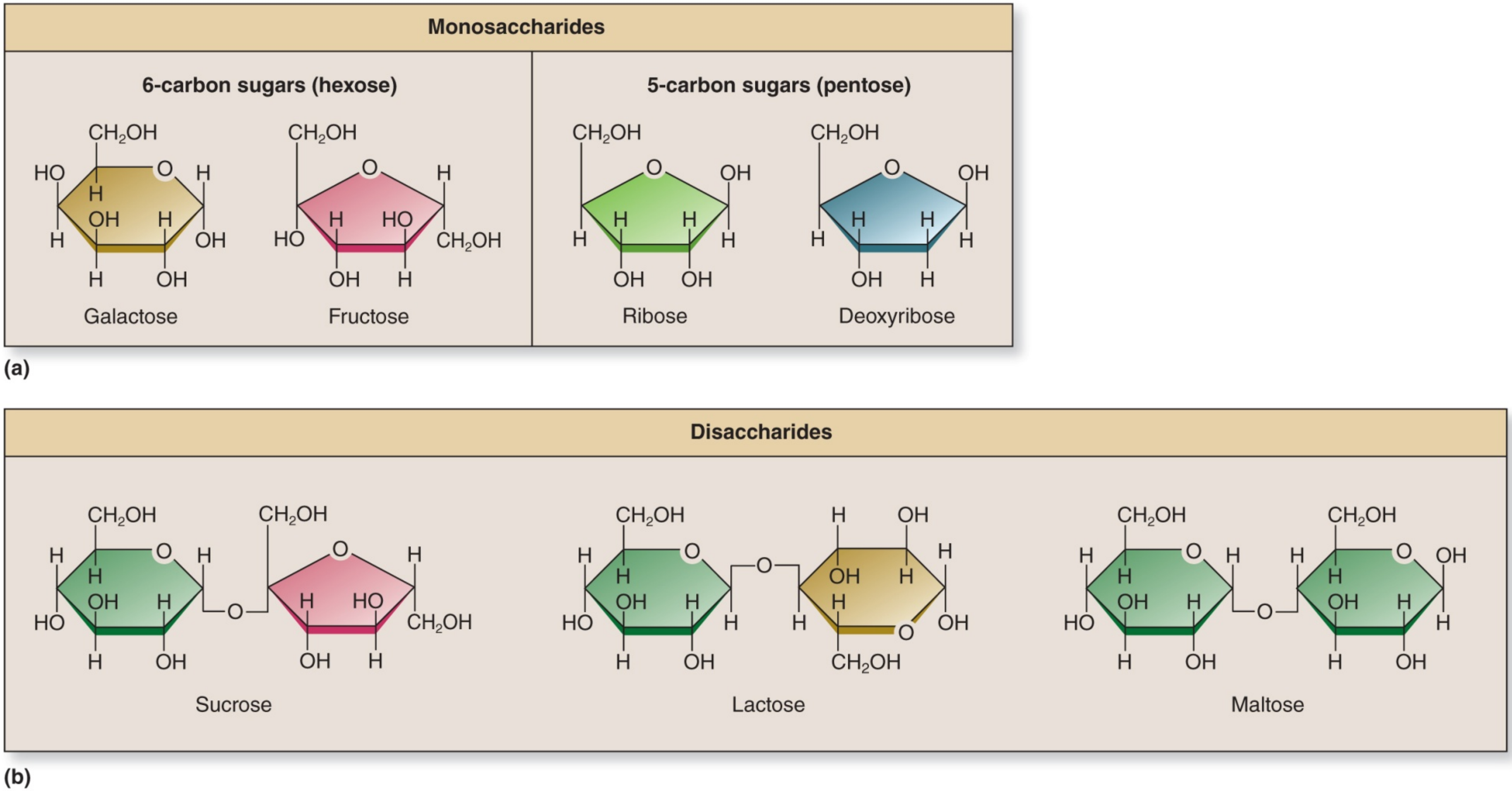


Figure 2.20

What did you learn?

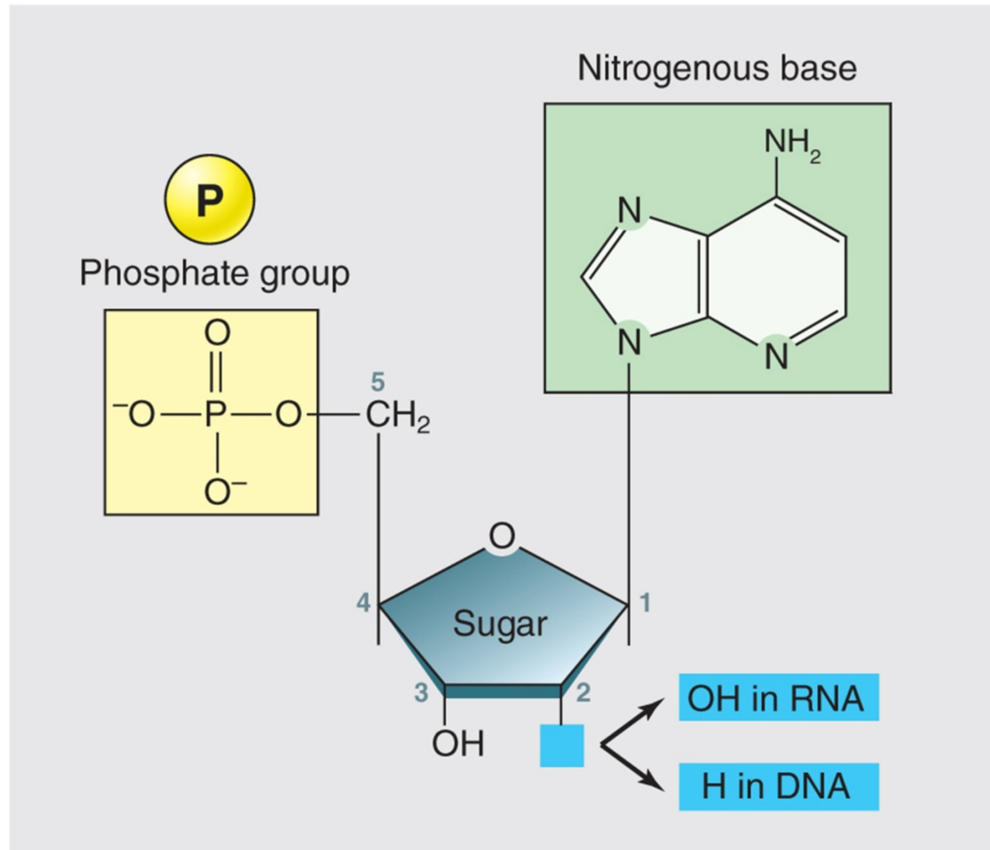
- What is the repeating monomer of glycogen?
- Where is glycogen stored in the body?
- Is galactose a monosaccharide, disaccharide or polysaccharide?
- Is sucrose a monosaccharide, disaccharide or polysaccharide?

2.7d Nucleic Acids

- **Nucleic acids**
  - Store and transfer genetic information
  - Two classes of nucleic acid
    - Deoxyribonucleic acid (DNA)
    - Ribonucleic acid (RNA)
      - Both are polymers composed of **nucleotide** monomers
      - Monomers are linked covalently through **phosphodiester bonds**

Nucleotide Monomer

- **3 components**
  1. **Sugar**
     – Five-carbon pentose
  2. **Phosphate group**
     – Attached at carbon 5 of sugar
  3. **Nitrogenous base**
     – Attached to same sugar at carbon one
     – Single-ring or double-ring structure



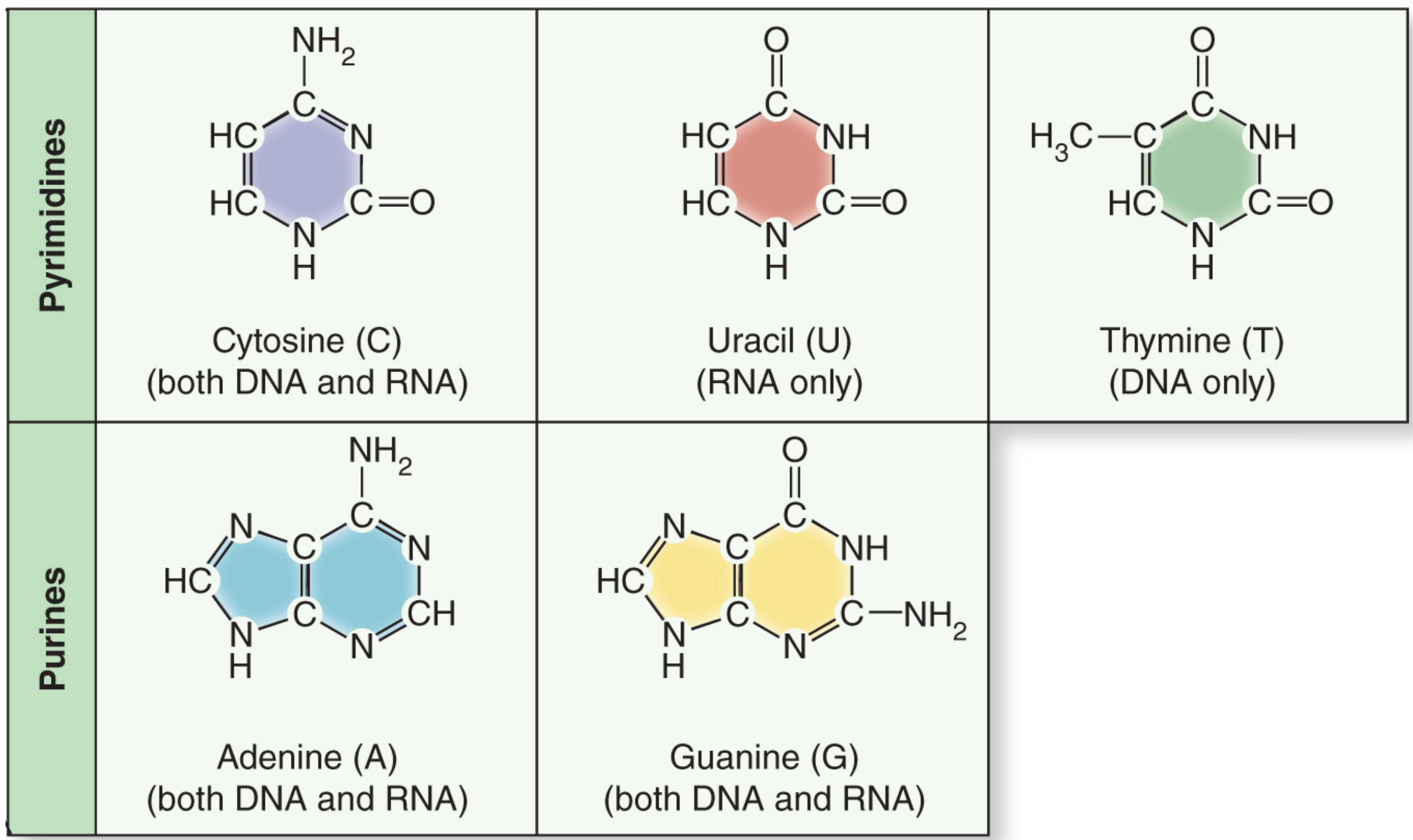
(a) Nucleotide monomer

Figure 2.21a

2.7d Nucleic Acids

- Five types of nitrogenous bases
  - **Pyrimidines—**single-ring bases
    - Cytosine
    - Uracil
    - Thymine
  - **Purines—**double-ring bases
    - Adenine
    - Guanine
- Nitrogenous bases within either group differ in functional groups attached to ring

Nitrogenous Bases



(b) Nitrogenous bases

Figure 2.21b

2.7d Nucleic Acids

- **Deoxyribonucleic acid (DNA)**
  - Double-stranded nucleic acid
  - Located in chromosomes in nucleus and in mitochondria
  - Deoxyribose sugar, phosphate, and one of four nitrogenous bases
    - Adenine, guanine, cystosine, thymine
    - No uracil
  - Double strands held together by hydrogen bonds
    - Form between complementary bases
      - Thymine paired with adenine
      - Guanine paired with cytosine

2.7d Nucleic Acids

- **Ribonucleic acid (RNA)**
  - Single-stranded nucleic acid
  - Located in nucleus and in cytoplasm of cell
  - Ribose sugar, phosphate, and one of four nitrogenous bases
    - Adenine
    - Guanine
    - Cystosine
    - Uracil
    - No thymine

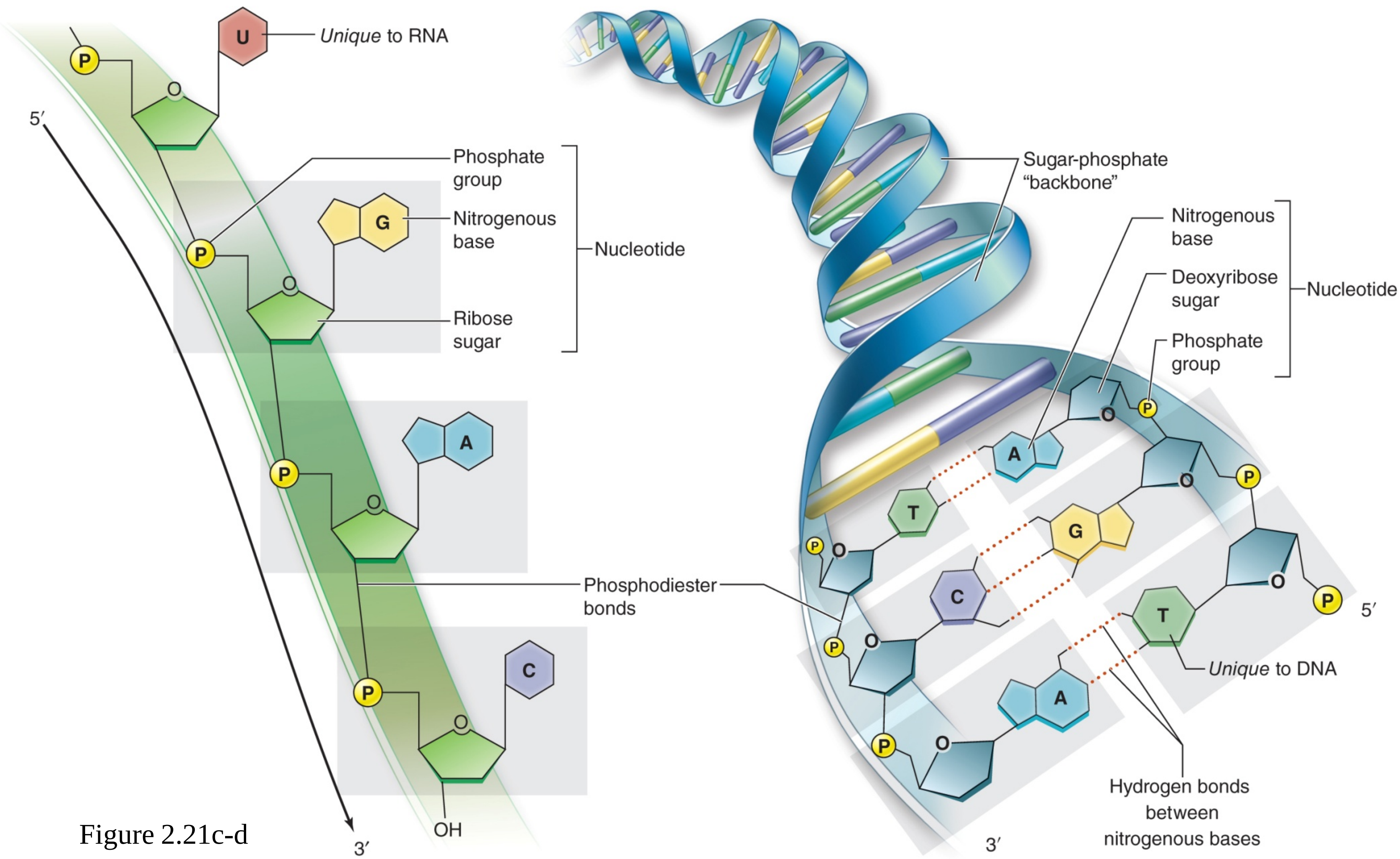


Figure 2.21c-d

(c) RNA (single-stranded)

(d) DNA (double-stranded)

2.7d Nucleic Acids

- **Adenosine triphosphate (ATP)**
  - Nucleotide composed of nitrogenous base adenine, a ribose sugar, and three phosphate groups
  - Central molecule in transfer of chemical energy within cell
  - Covalent bonds between last two phosphate groups are unique, energy rich
  - Release energy when broken
- Important nucleotide-containing molecules
  - Nicotinamide adenine dinucleotide
  - Flavin adenine dinucleotide
  - Both participate in production of ATP

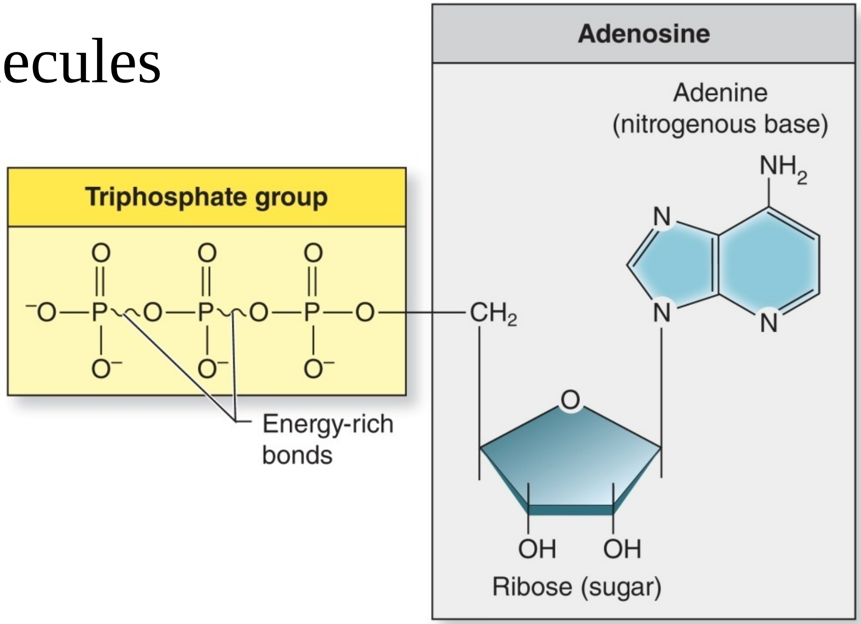


Figure 2.22

What did you learn?

- What are the structural differences between RNA and DNA?

2.7e Proteins

- Functions of proteins
  - Serve as *catalysts* (enzymes) in metabolic reactions
  - Act in *defense*
  - Aid in *transport*
  - Contribute to structural *support*
  - Cause *movement*
  - Perform *regulation*
  - Provide *storage*

2.7e Proteins

- General protein structure
  - One or more strands of **amino acid** monomers
  - 20 different amino acids found in living organisms
  - Each has an amine and a carboxyl functional group
    - Both covalently linked to same carbon atom
  - Carbon also covalently bonded to a hydrogen and different side chain structures
    - R groups, which distinguish different amino acids from one another

2.7e Proteins

- General protein structure (*continued*)
  - Amino acids are covalently linked by **peptide bonds**
    - Formed during dehydration synthesis reaction between amine group of one amino acid and the carboxylic group of another
    - —H lost from the amine group
    - — OH lost from the carboxylic acid of another amino acid
  - **N-terminal end** has free amine group
  - **C-terminal end** has free carboxyl group

Proteins

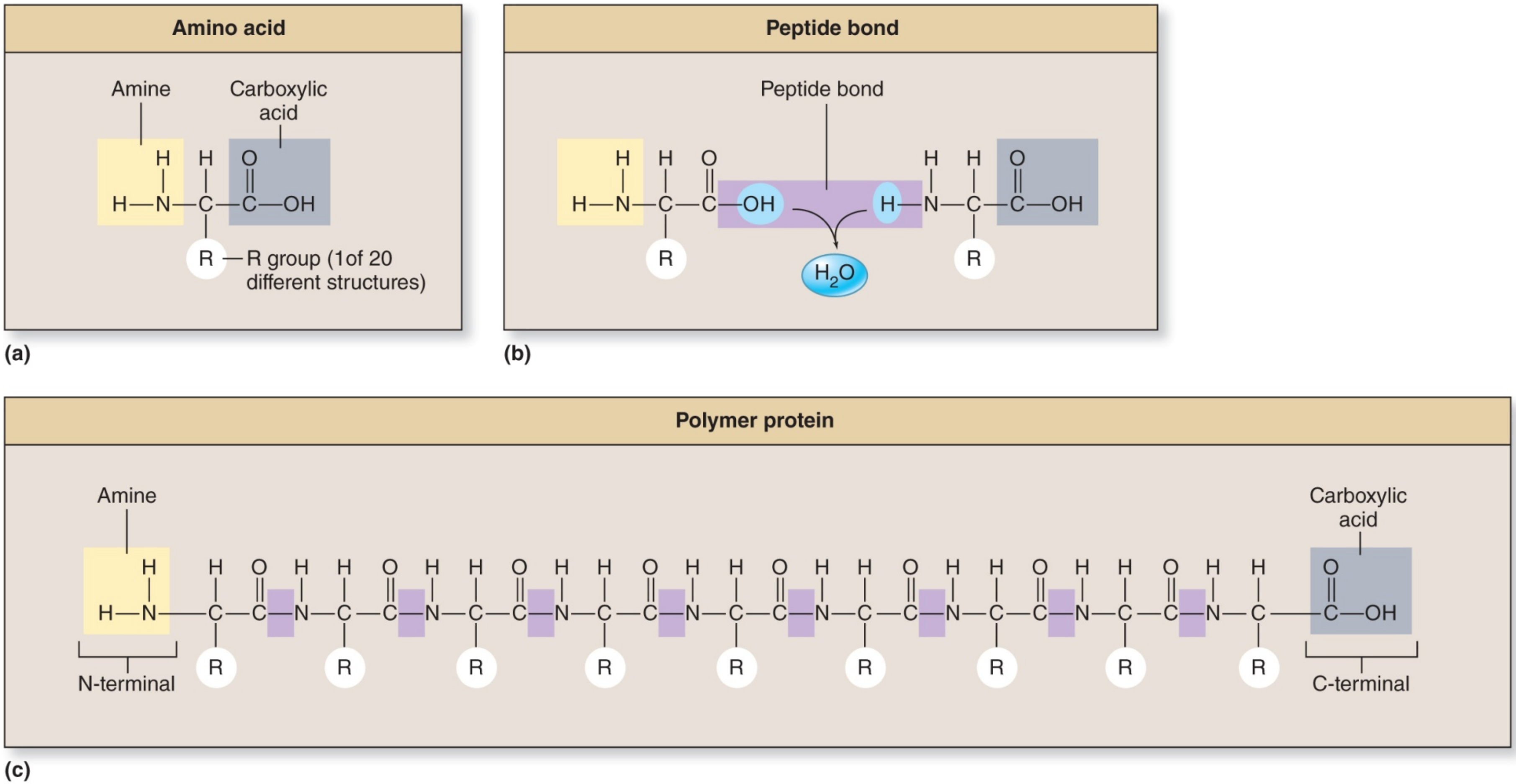


Figure 2.23

2.7e Proteins

- General protein structure (*continued*)
  - Strands of amino acids
    - **Oligopeptide:** between 3 and 20 amino acids
    - **Polypeptide:** between 21 and 199 amino acids
    - **Protein:** more than 200 amino acids
- **Glycoproteins** are proteins with carbohydrate attached
  - E.g., glycoproteins on erythrocytes determining ABO blood groups
- Summary
  - Lipids, carbohydrates, nucleic acids, proteins are four major classes of organic biomolecules that compose human body

What did you learn?

- What are the monomers of proteins?
- What is the name of the bond between them?

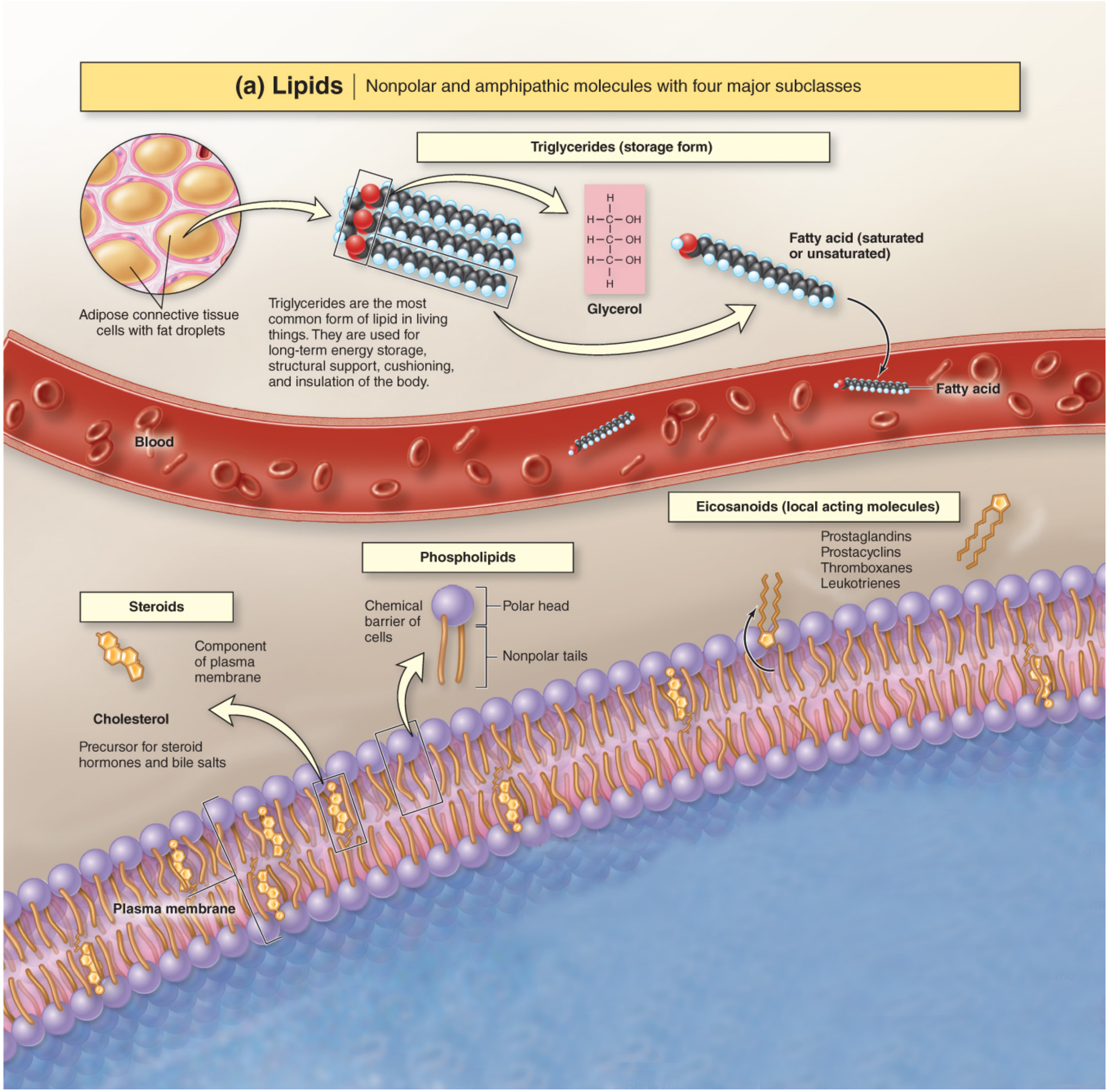
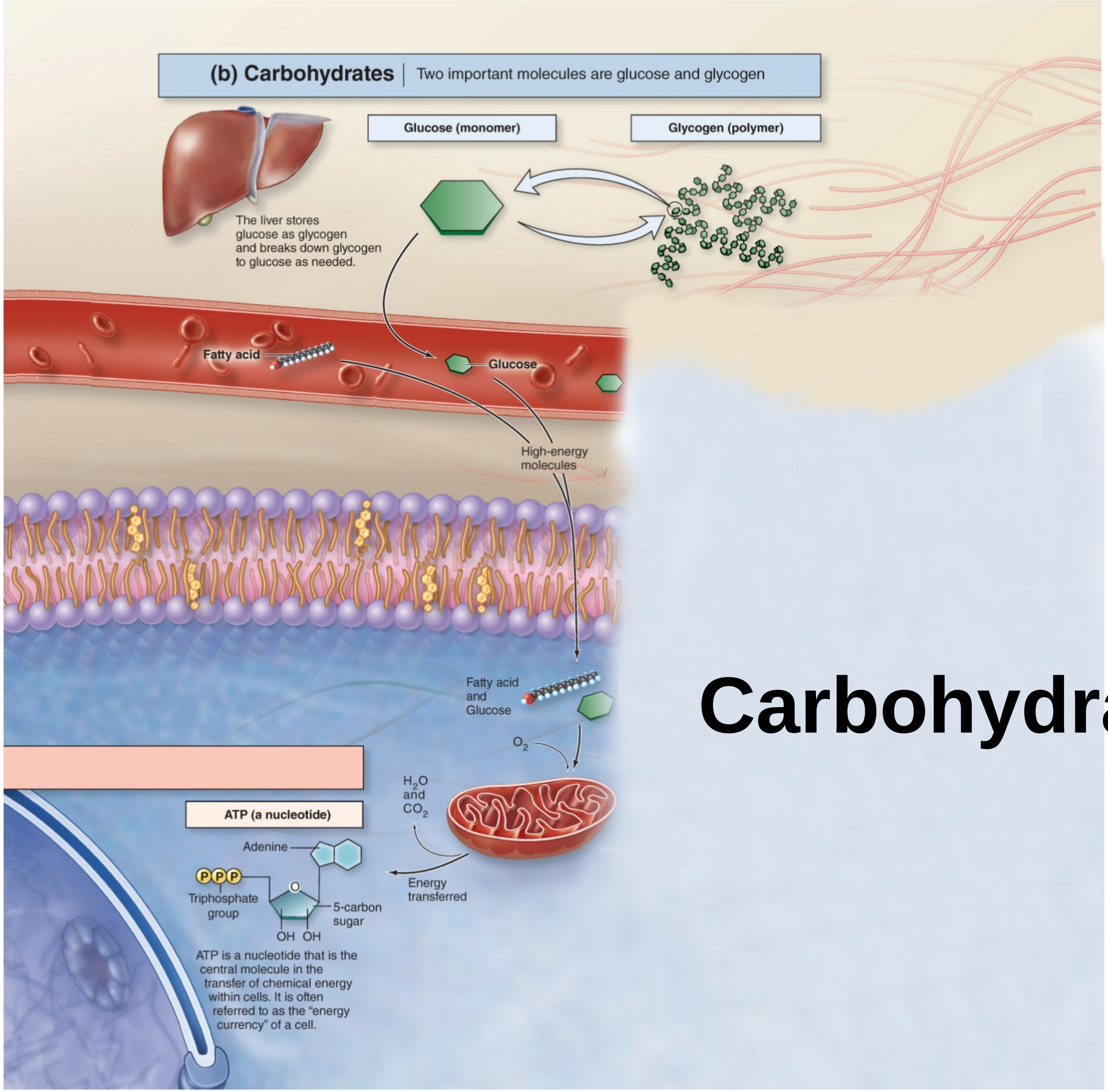

(a) Lipids
Nonpolar and amphipathic molecules with four major subclasses
Triglycerides (storage form)
Adipose connective tissue cells with fat droplets
Triglycerides are the most common form of lipid in living things. They are used for long-term energy storage, structural support, cushioning, and insulation of the body.
H
H–C–OH
H–C–OH
H–C–OH
H
Glycerol
Fatty acid (saturated or unsaturated)
Fatty acid
Blood
Eicosanoids (local acting molecules)
Prostaglandins
Prostacyclins
Thromboxanes
Leukotrienes
Phospholipids
Chemical barrier of cells
Polar head
Nonpolar tails
Steroids
Component of plasma membrane
Cholesterol
Precursor for steroid hormones and bile salts
Plasma membrane

Figure 2.24a



Carbohydrates

Figure 2.24b

Nucleic Acids and Nucleotides

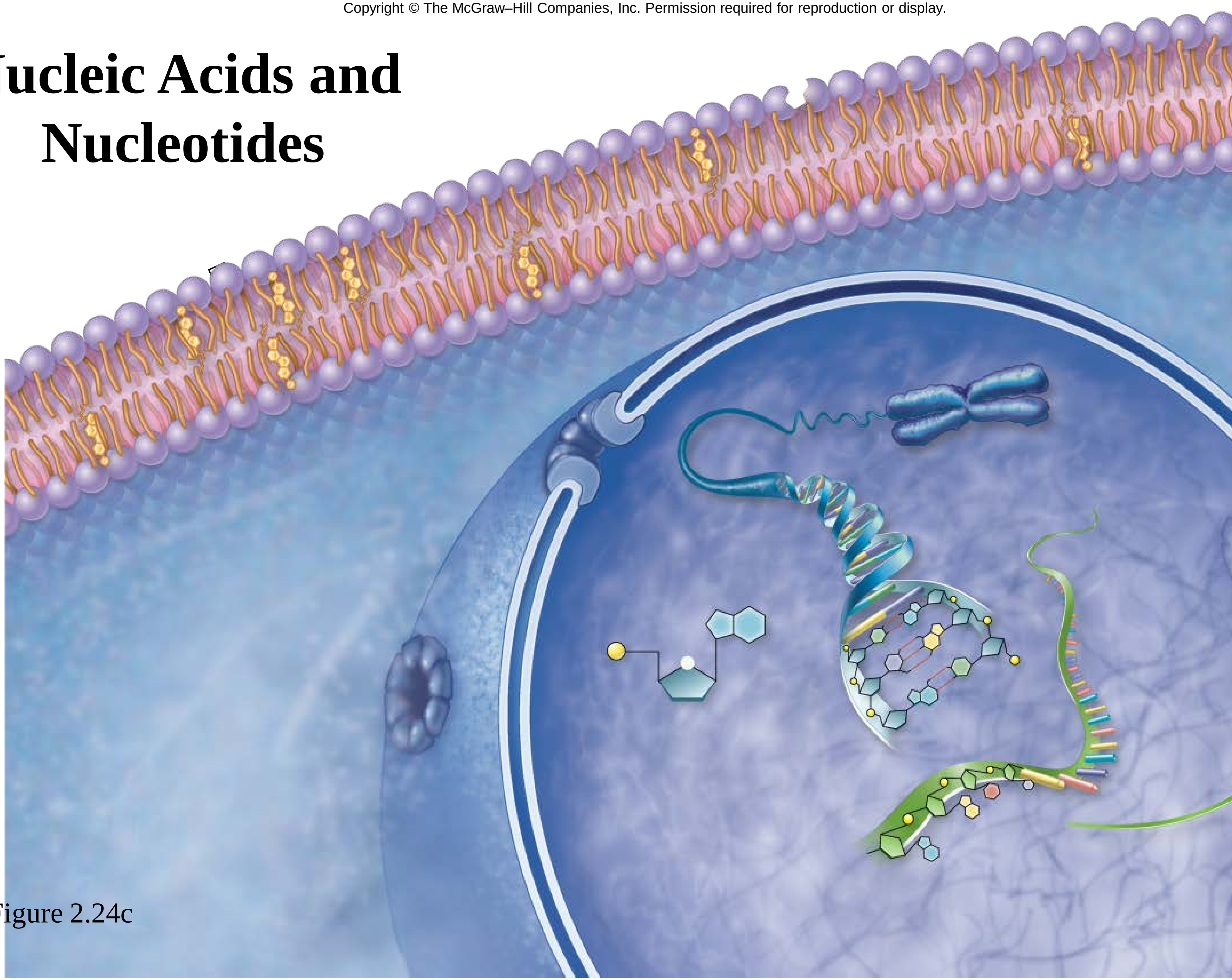


Figure 2.24c

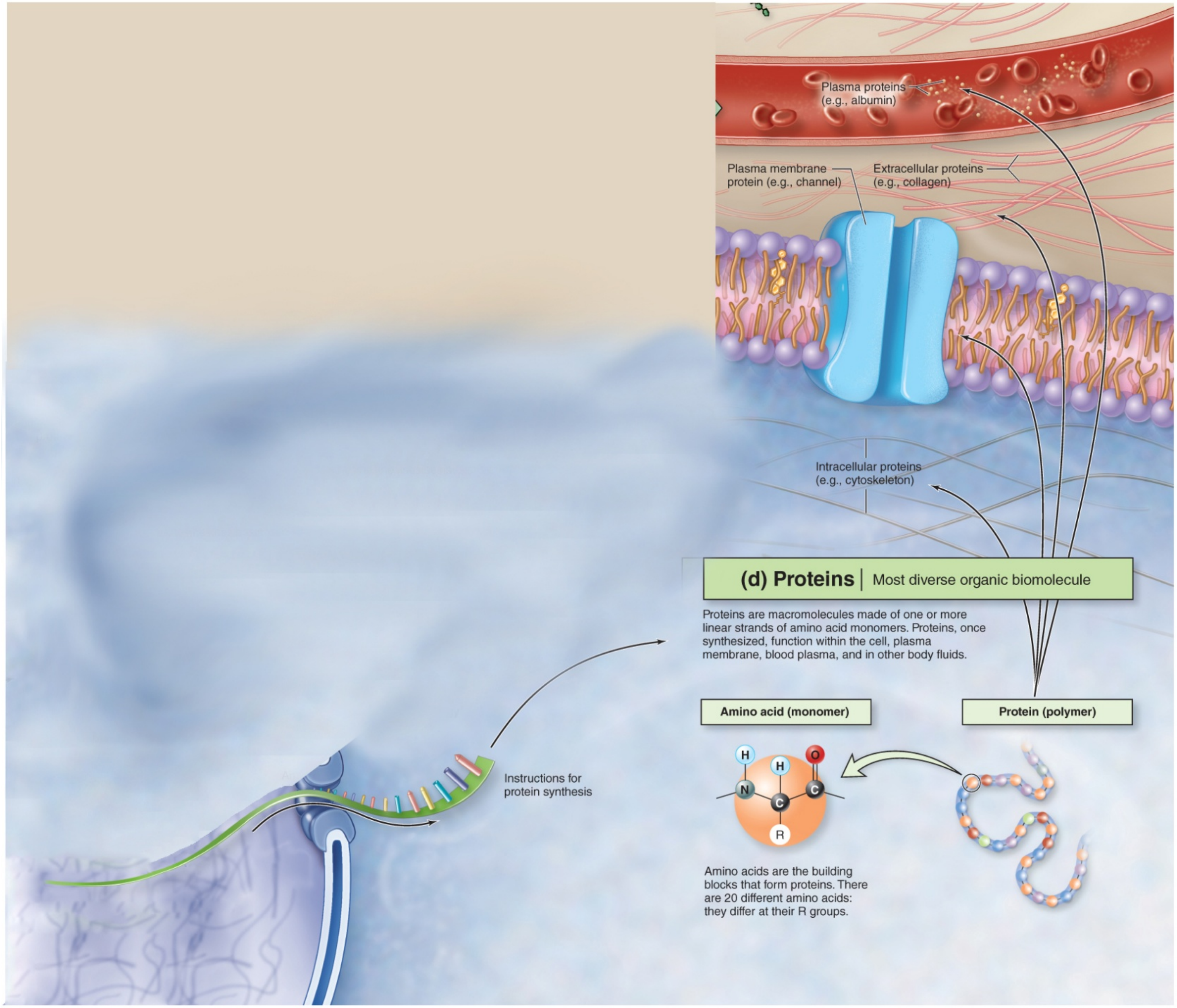

Plasma proteins
(e.g., albumin)
Plasma membrane protein (e.g., channel)
Extracellular proteins (e.g., collagen)
Intracellular proteins (e.g., cytoskeleton)
(d) Proteins | Most diverse organic biomolecule
Proteins are macromolecules made of one or more linear strands of amino acid monomers. Proteins, once synthesized, function within the cell, plasma membrane, blood plasma, and in other body fluids.
Amino acid (monomer)
Protein (polymer)
Instructions for protein synthesis
H
H
O
N
C
C
R
Amino acids are the building blocks that form proteins. There are 20 different amino acids: they differ at their R groups.

Figure 2.24d

2.8 Protein Structure

Learning Objectives:

1. Name the categories of amino acids.
2. Distinguish between nonpolar, polar, and charged amino acids.
3. Give examples of amino acids with special characteristics.

2.8 Protein Structure (*continued*)

Learning Objectives:

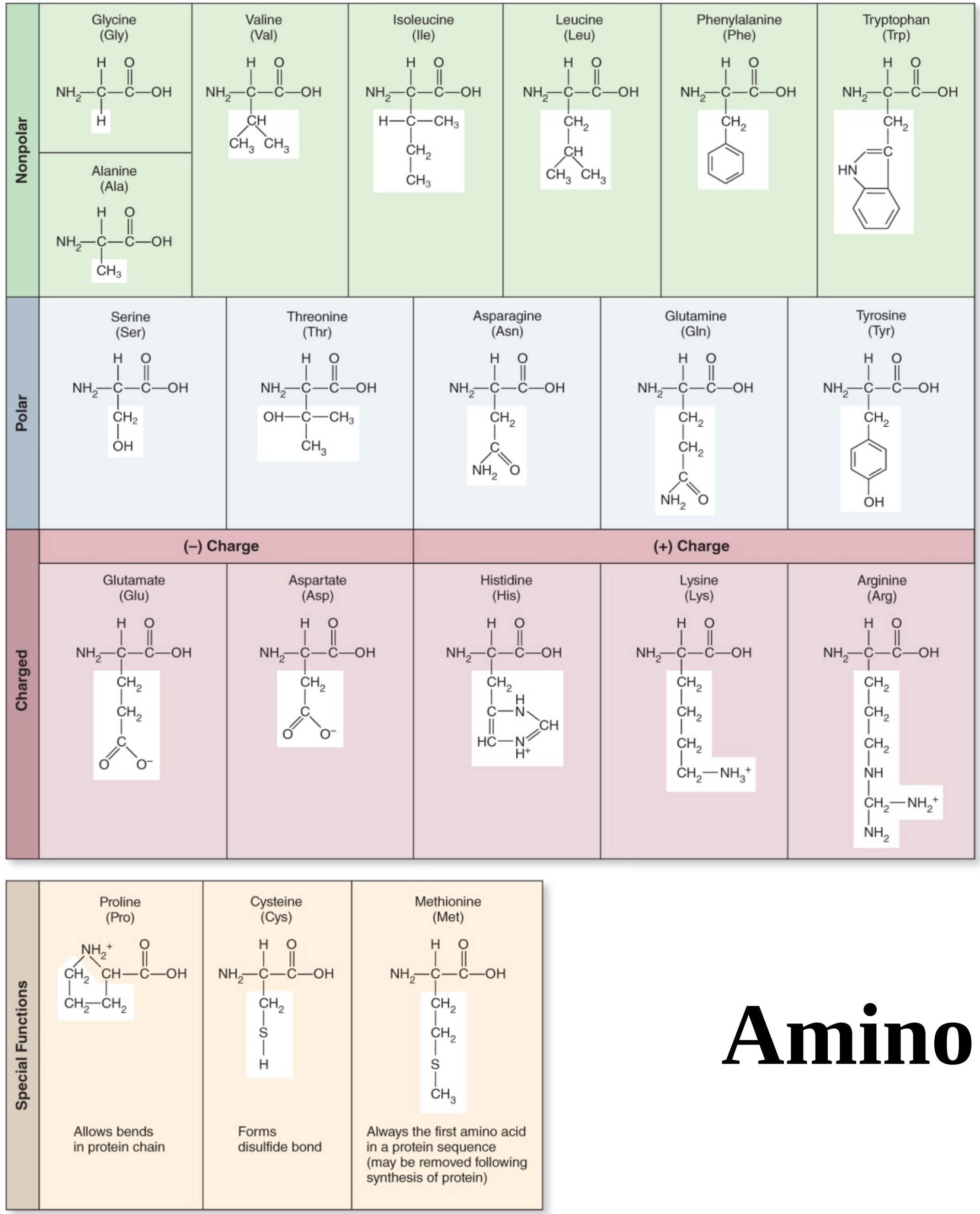
4. Describe the different types of intramolecular attractions that participate in both the folding of a protein and in maintaining its three-dimensional shape.
5. Distinguish between the four structural hierarchy levels of proteins.
6. Explain what is meant by denaturation and list factors that can cause it.

2.8a Categories of Amino Acids

- Organized into groups based on their R group
  - **Nonpolar amino acids**
    - Contain R groups with hydrogen or hydrocarbons
    - Group with other nonpolar amino acids by hydrophobic interactions
  - **Polar amino acids**
    - Contain R groups with other elements besides hydrogen or hydrocarbons
    - Form interactions with other polar amino acids and with water

2.8a Categories of Amino Acids

- Organized into groups based on their R group (*continued*)
  - **Charged amino acids**
    - Contain R groups with a negative or a positive charge
    - Form ionic bonds between negatively and positively charged amino acids
    - Hydrophilic
  - **Amino acids with special functions**
    - Proline can cause a bend in the protein chain
    - Cysteine can form **disulfide bond**
    - Methionine, first amino during protein synthesis



Amino Acids

Figure 2.25

2.8b Amino Acid Sequence and Protein Conformation

- **Primary structure—**linear sequence of amino acids

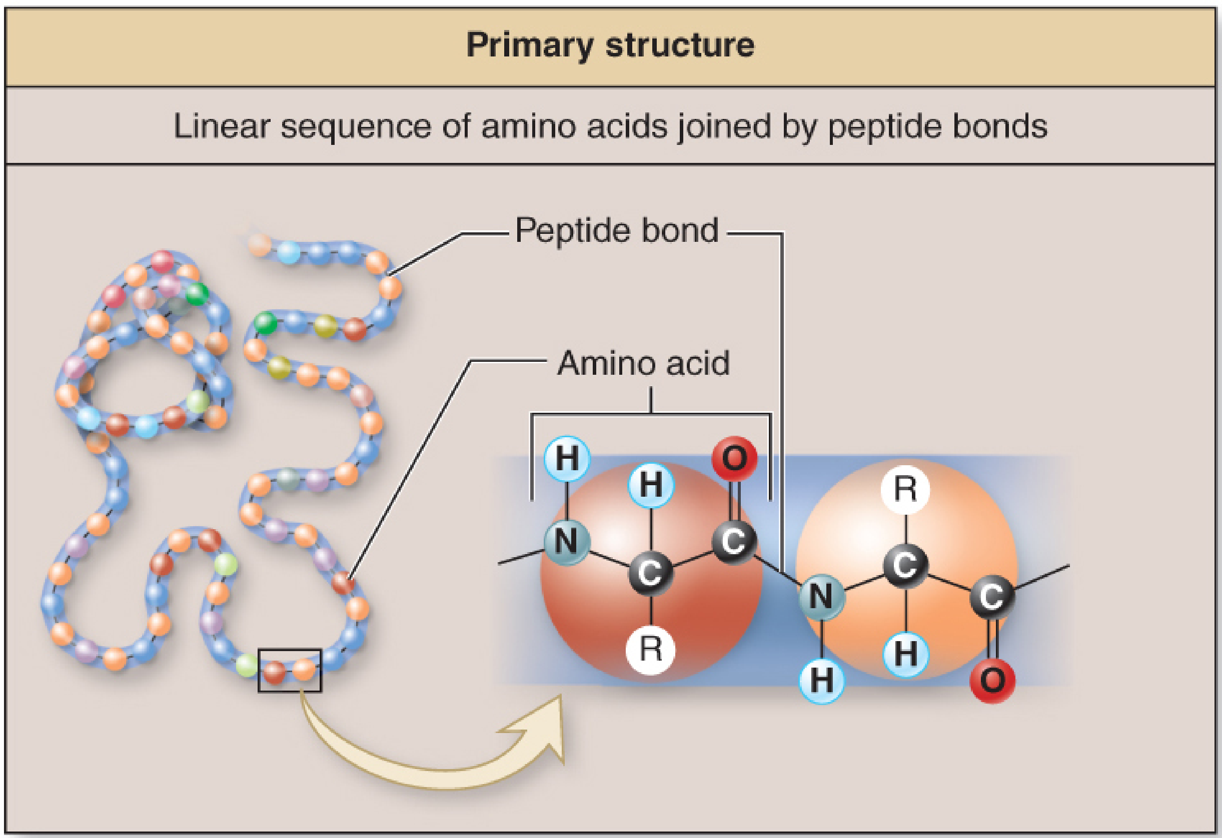


Figure 2.26a

(a)

2.8b Amino Acid Sequence and Protein Conformation

- **Conformation—**three-dimensional shape of the protein
  - Crucial for protein function
  - Levels of organization beyond primary structure
  - Arrangements dependent upon intramolecular attractions between amino acids
  - Obtained through folding with help of specialized proteins, *chaperones*

2.8b Amino Acid Sequence and Protein Conformation

- Intramolecular interactions
  - Hydrophobic interactions with nonpolar amino acids farther from water
  - Hydrogen bonds between polar R groups, between amine and carboxylic acid groups
  - Ionic bonds between negative and positive R groups
  - Disulfide bonds between cysteine amino acids

2.8b Amino Acid Sequence and Protein Conformation

- **Secondary structures**
  - Patterns that may repeat several times
  - Confer unique characteristics
  - Two types
    - **Alpha helix**—spiral coil
      - Elasticity to fibrous proteins (e.g., skin and hair)
    - **Beta sheet**—planar pleat arrangement
      - Flexibility to globular proteins (e.g., enzymes)

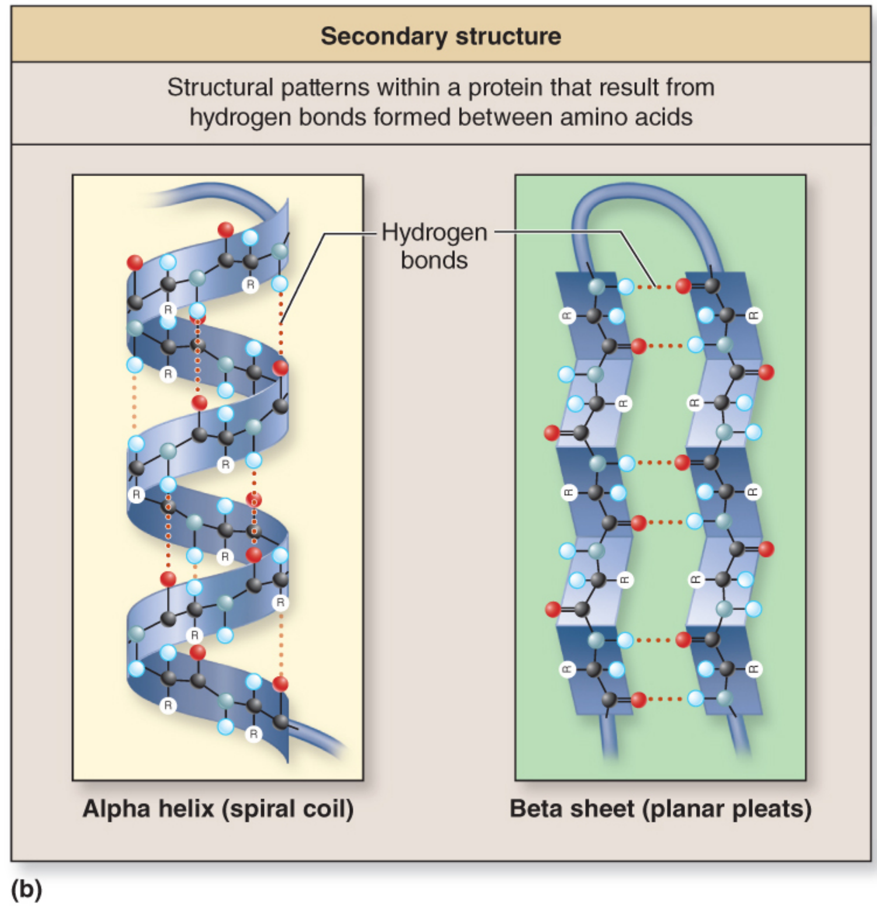


Figure 2.26b

2.8b Amino Acid Sequence and Protein Conformation

- **Tertiary structure**
  - Three-dimensional shape of poly-peptide chain
  - Two categories
    - **Globular proteins** fold into compact shape
    - **Fibrous proteins** are extended linear molecules

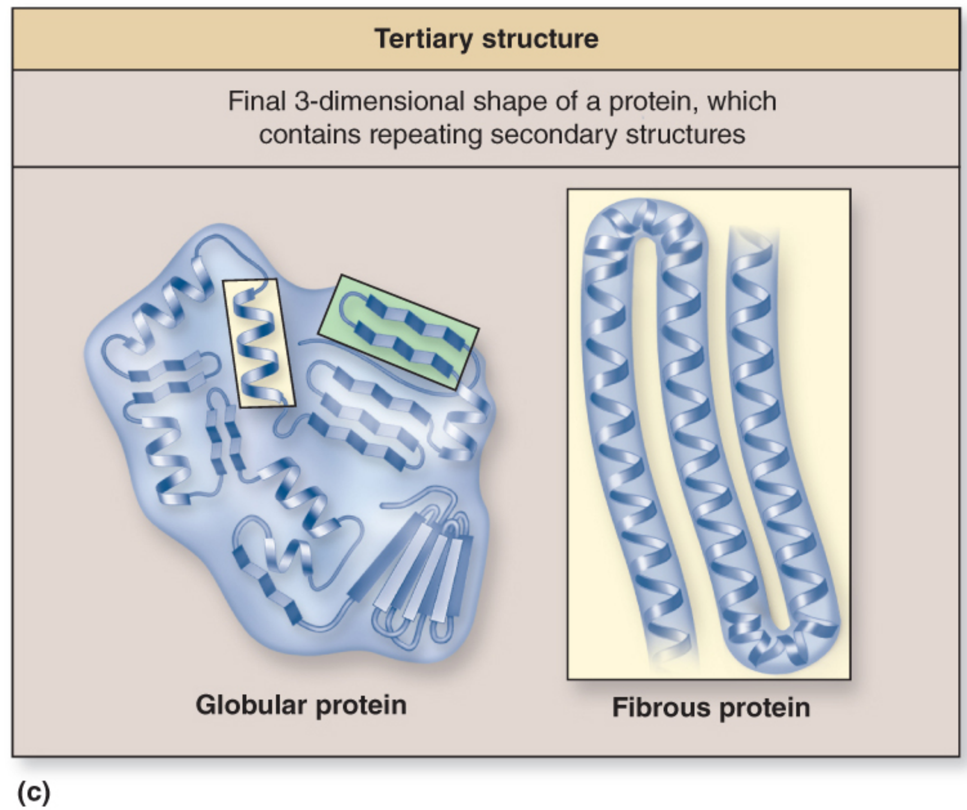


Figure 2.26c

2.8b Amino Acid Sequence and Protein Conformation

- **Quaternary structure**
  - Present in proteins with two or more polypeptide chains
  - E.g., hemoglobin with its four polypeptide chains

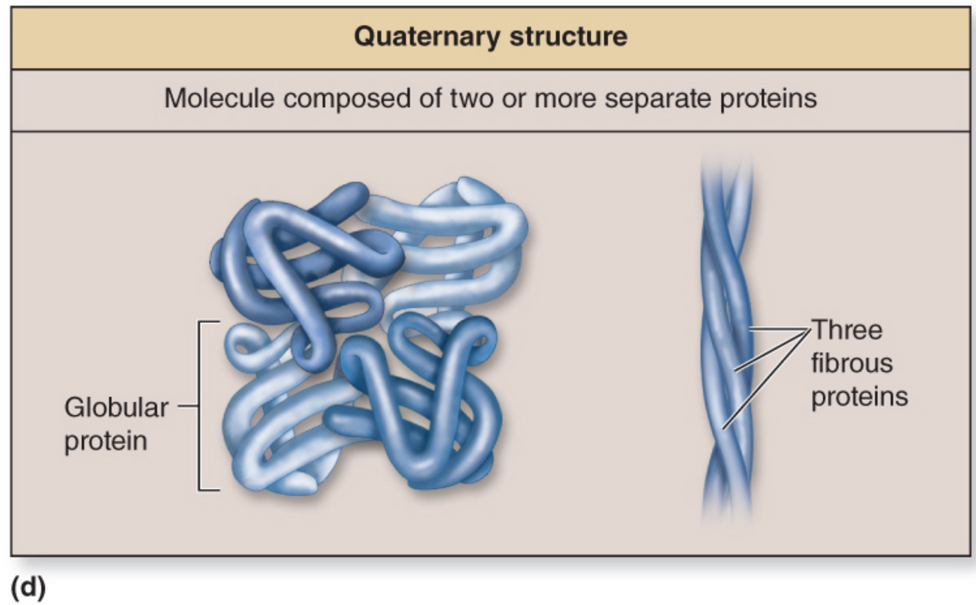


Figure 2.26d

2.8b Amino Acid Sequence and Protein Conformation

- **Prosthetic groups**
  - Nonprotein structures covalently bonded to protein
    - E.g., lipid heme group in hemoglobin protein
- **Denaturation**
  - Conformational change to a protein
  - Disturbs protein activity
  - Usually irreversible
  - May occur due to increased temperature or in response to changes in pH

2.8b Amino Acid Sequence and Protein Conformation

- Other causes of denaturation
  - pH changes
    - o Interfere with electrostatic interactions and some other intramolecular bonds
    - o Changes in blood pH can be lethal

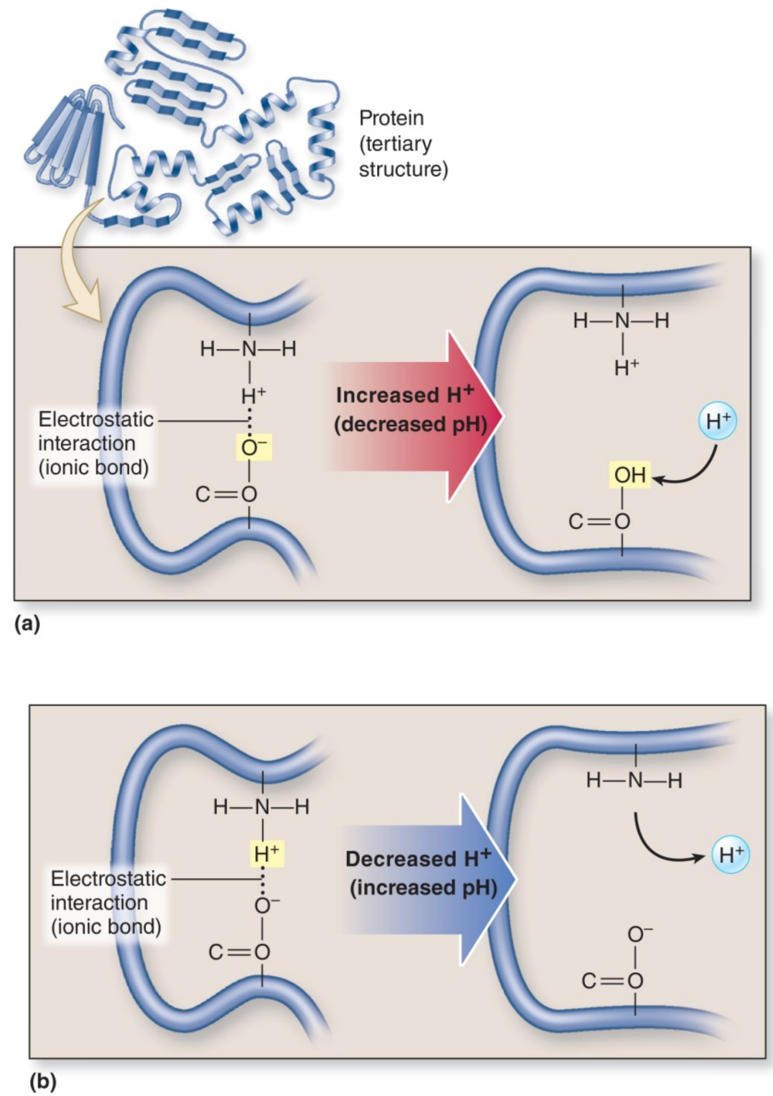


Figure 2.27

What did you learn?

- Distinguish between the primary, secondary, tertiary, and quaternary levels of protein organization?