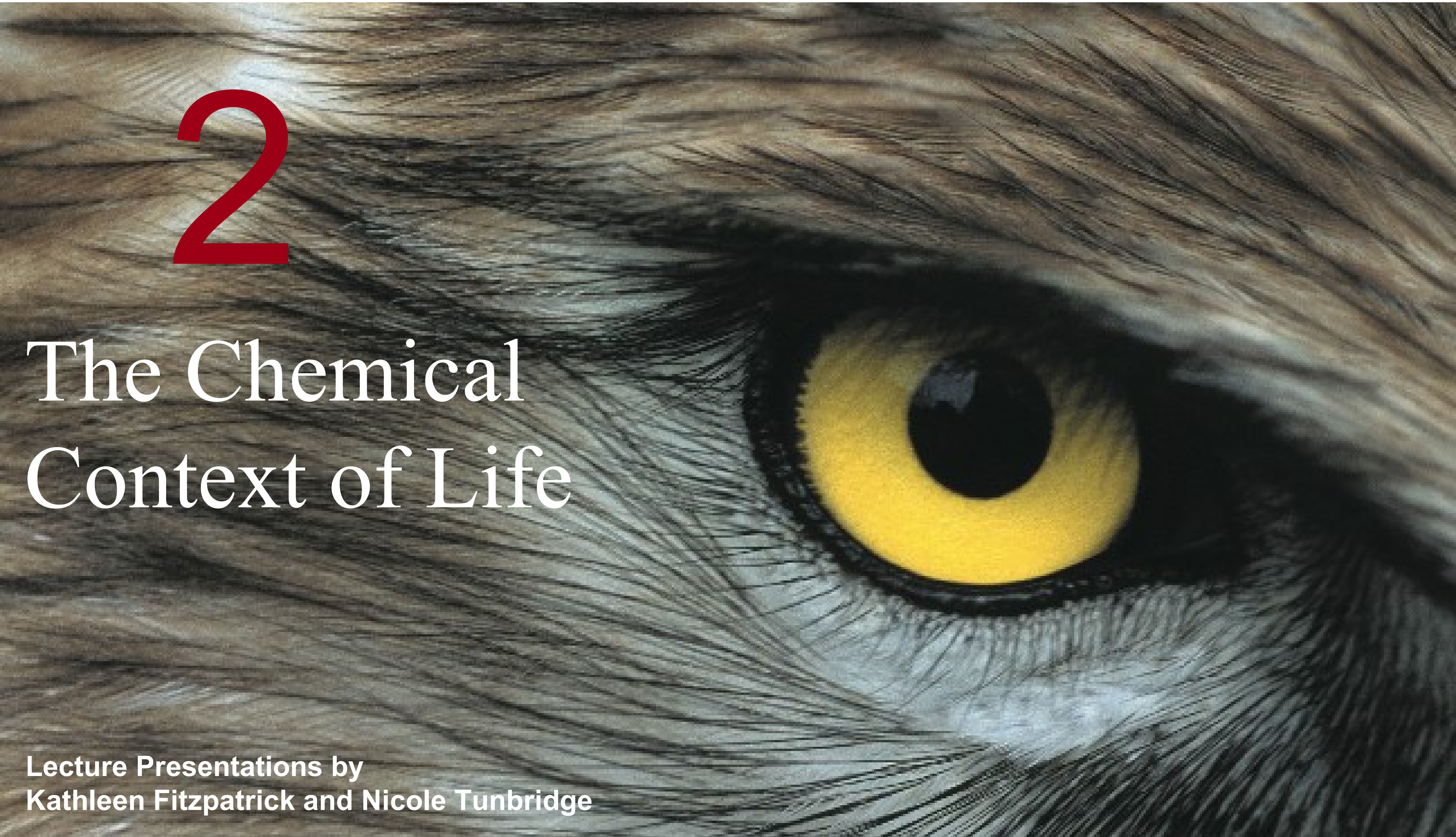
CAMPBELL BIOLOGY IN FOCUS

Urry • Cain • Wasserman • Minorsky • Jackson • Reece



2

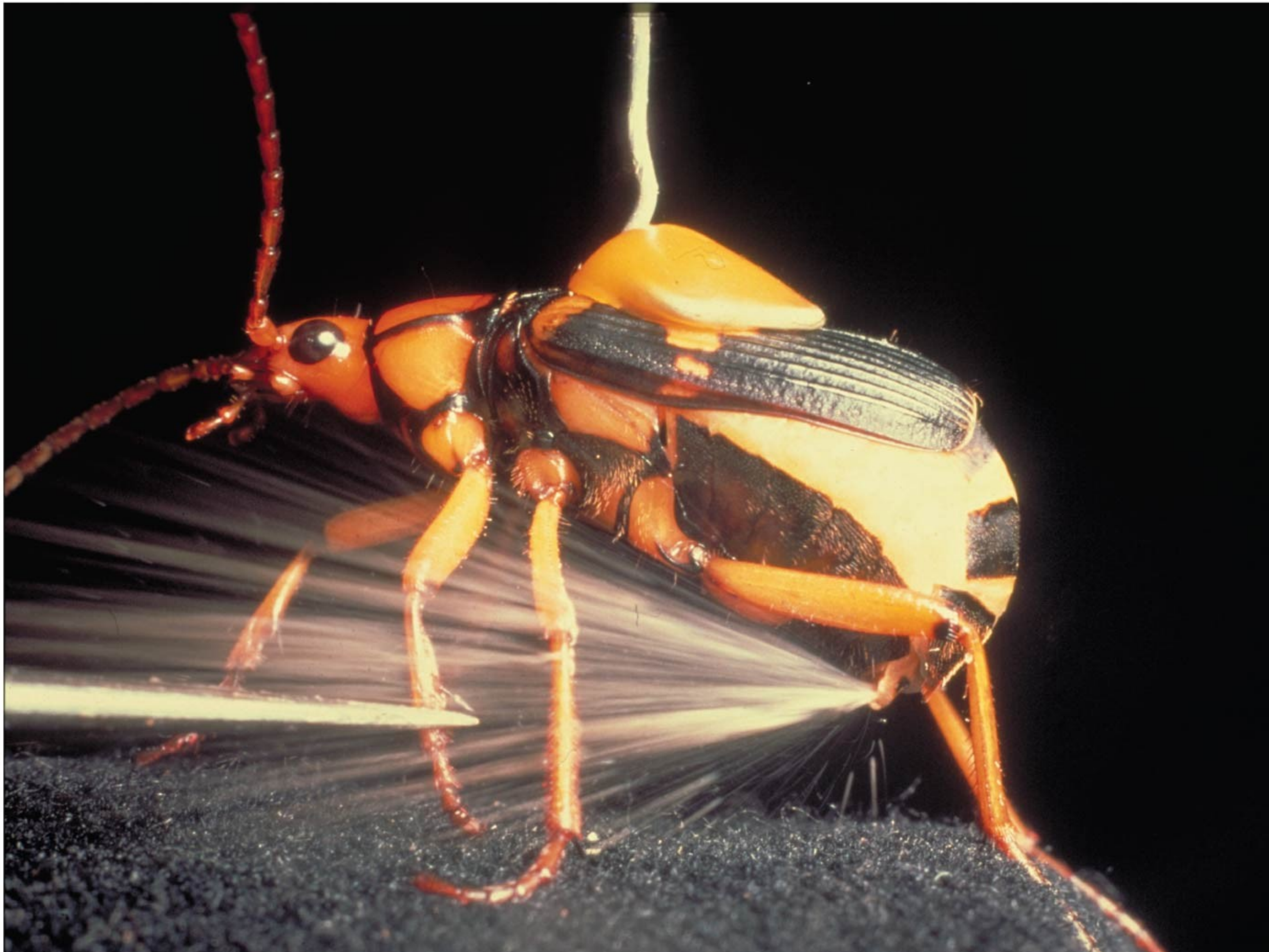
The Chemical Context of Life

Lecture Presentations by
Kathleen Fitzpatrick and Nicole Tunbridge

Overview: A Chemical Connection to Biology

- Biology is a multidisciplinary science
- Living organisms are subject to basic laws of physics and chemistry

Figure 2.1



Concept 2.1: Matter consists of chemical elements in pure form and in combinations called compounds

- Organisms are composed of **matter**
- Matter is anything that takes up space and has mass

Elements and Compounds

- Matter is made up of elements
- An **element** is a substance that cannot be broken down to other substances by chemical reactions
- A **compound** is a substance consisting of two or more elements in a fixed ratio
- A compound has emergent properties, characteristics different from those of its elements

Figure 2.2

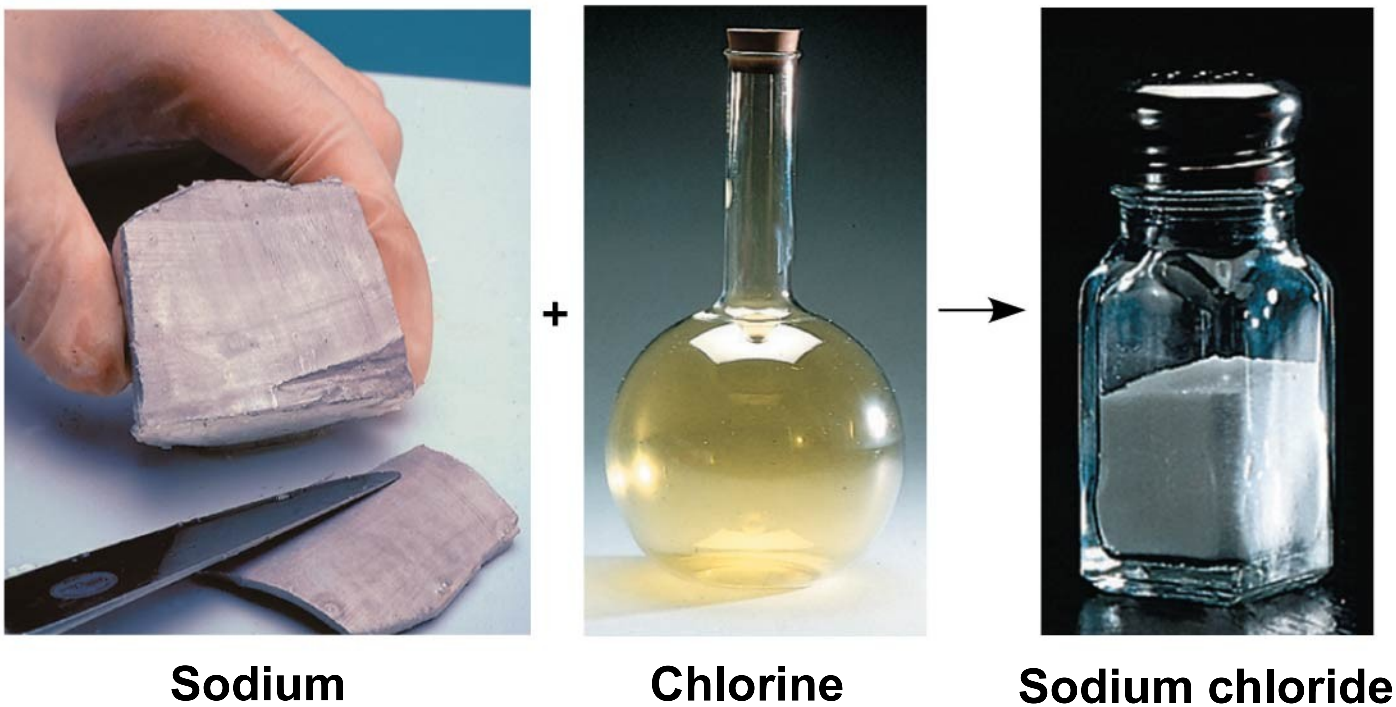


Figure 2.2a

Sodium

Figure 2.2b

Chlorine

Figure 2.2c

Sodium chloride

The Elements of Life

- Of 92 natural elements, about 20–25% are **essential elements**, needed by an organism to live a healthy life and reproduce
- **Trace elements** are required in only minute quantities
- For example, in vertebrates, iodine (I) is required for normal activity of the thyroid gland
- In humans, an iodine deficiency can cause goiter

Evolution of Tolerance to Toxic Elements

- Some naturally occurring elements are toxic to organisms
- In humans, arsenic is linked to many diseases and can be lethal
- Some species have become adapted to environments containing elements that are usually toxic
  - For example, sunflower plants can take up lead, zinc, and other heavy metals in concentrations lethal to most organisms
  - Sunflower plants were used to detoxify contaminated soils after Hurricane Katrina

Concept 2.2: An element's properties depend on the structure of its atoms

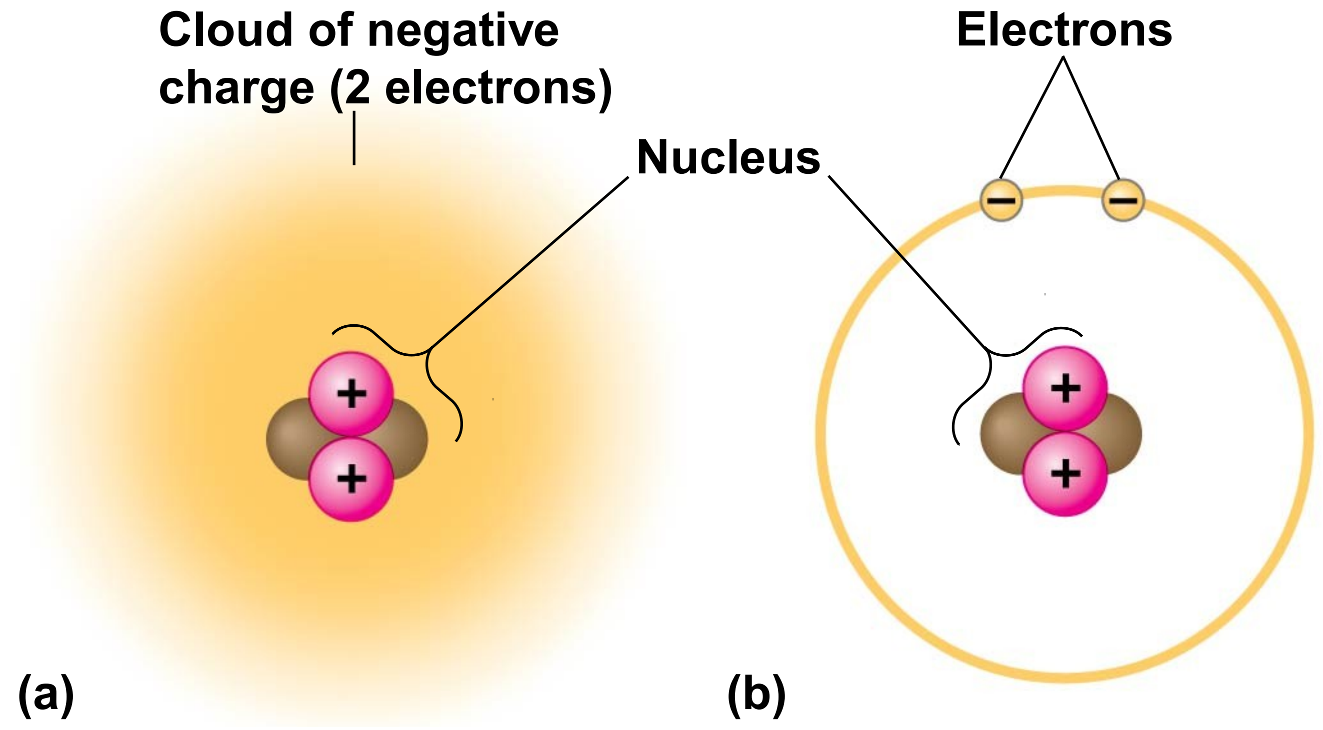
- Each element consists of a certain type of atom, different from the atoms of any other element
- An **atom** is the smallest unit of matter that still retains the properties of an element

Subatomic Particles

- Atoms are composed of smaller parts called subatomic particles
- Relevant subatomic particles include
  - **Neutrons** (no electrical charge)
  - **Protons** (positive charge)
  - **Electrons** (negative charge)

- Neutrons and protons form the **atomic nucleus**
- Electrons form a cloud around the nucleus
- Neutron mass and proton mass are almost identical and are measured in **daltons**

Figure 2.3



Atomic Number and Atomic Mass

- Atoms of the various elements differ in number of subatomic particles
- An element's **atomic number** is the number of protons in its nucleus
- An element's **mass number** is the sum of protons plus neutrons in the nucleus
- **Atomic mass**, the atom's total mass, can be approximated by the mass number

Mass number = number of protons + neutrons
= 23 for sodium

$^{23}_{11}\text{Na}$

Atomic number = number of protons
= 11 for sodium

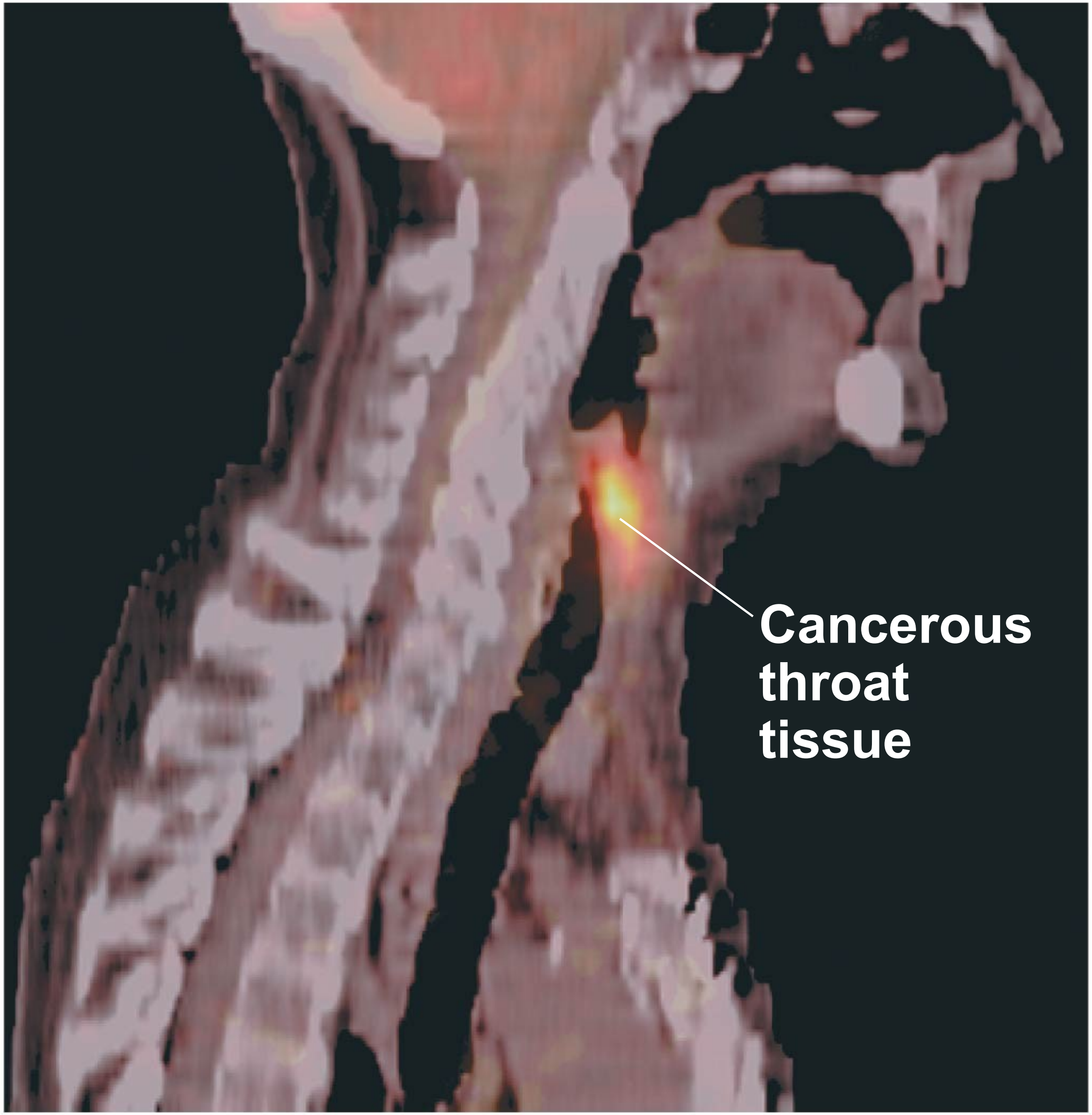
Because neutrons and protons each have a mass of approximately 1 dalton, we can estimate the **atomic mass** (total mass of one atom) of sodium as 23 daltons

Isotopes

- All atoms of an element have the same number of protons but may differ in number of neutrons
- **Isotopes** are two atoms of an element that differ in number of neutrons
- **Radioactive isotopes** decay spontaneously, giving off particles and energy

- Some applications of radioactive isotopes in biological research are
  - Dating fossils
  - Tracing atoms through metabolic processes
  - Diagnosing medical disorders

Figure 2.4

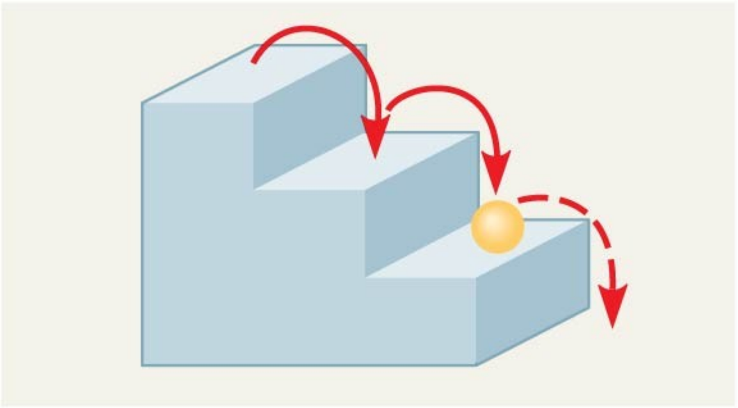


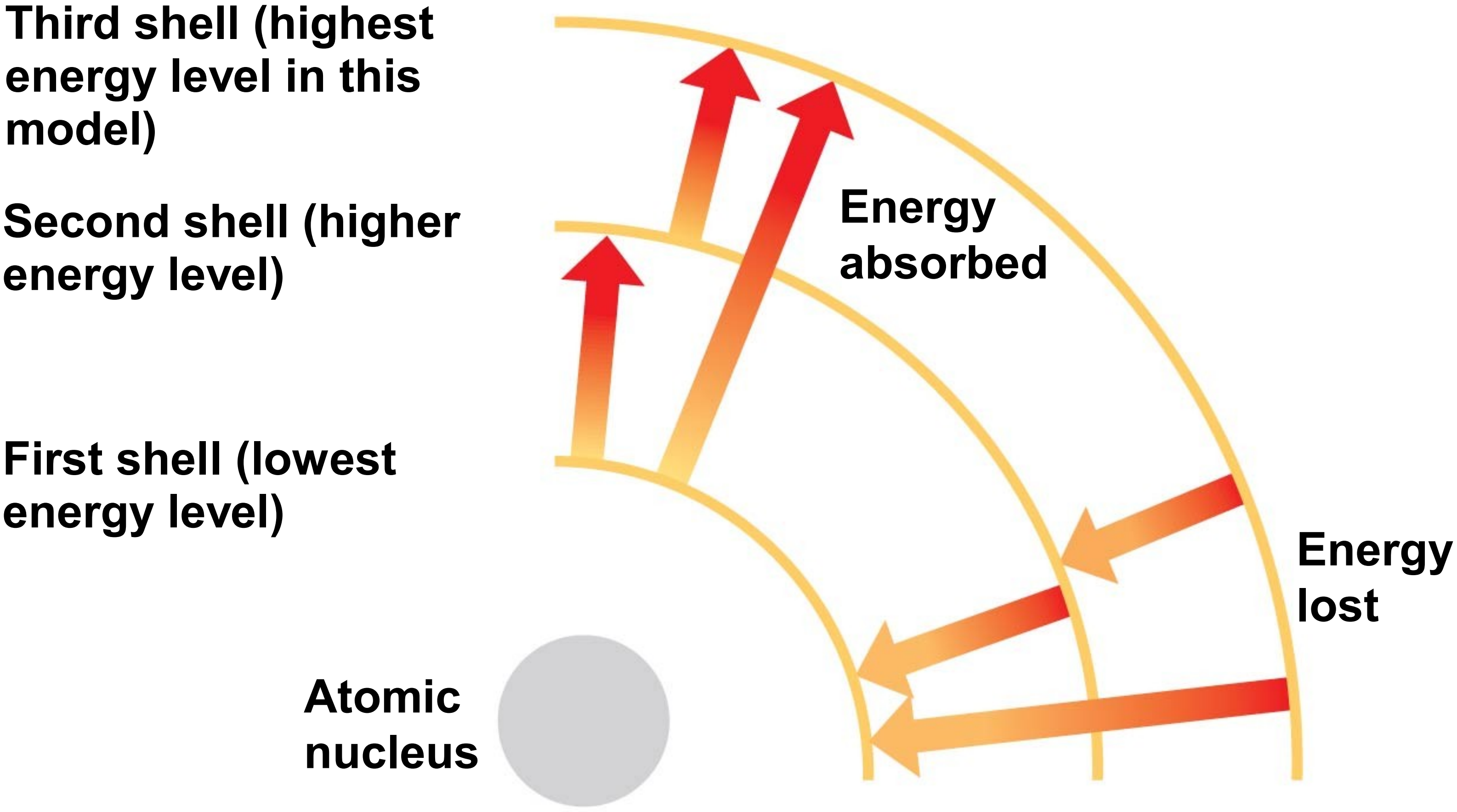
The Energy Levels of Electrons

- **Energy** is the capacity to cause change
- **Potential energy** is the energy that matter has because of its location or structure
- The electrons of an atom differ in their amounts of potential energy
- Changes in potential energy occur in steps of fixed amounts
- An electron’s state of potential energy is called its energy level, or **electron shell**

Figure 2.5

(a) A ball bouncing down a flight of stairs provides an analogy for energy levels of electrons.

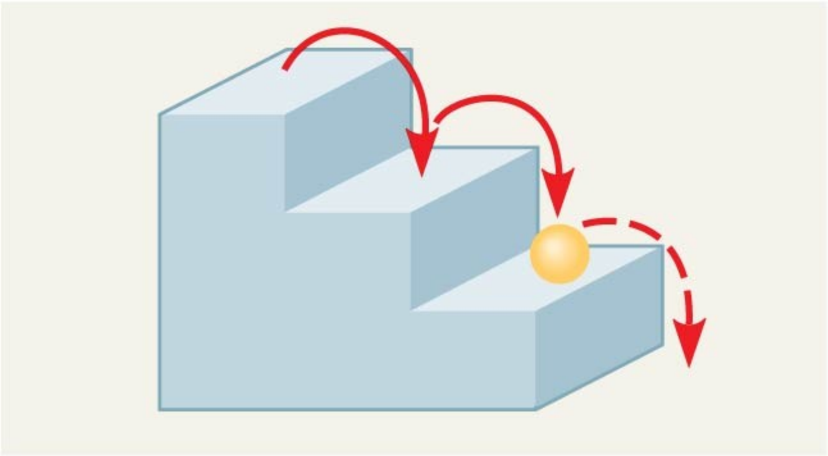




(b)

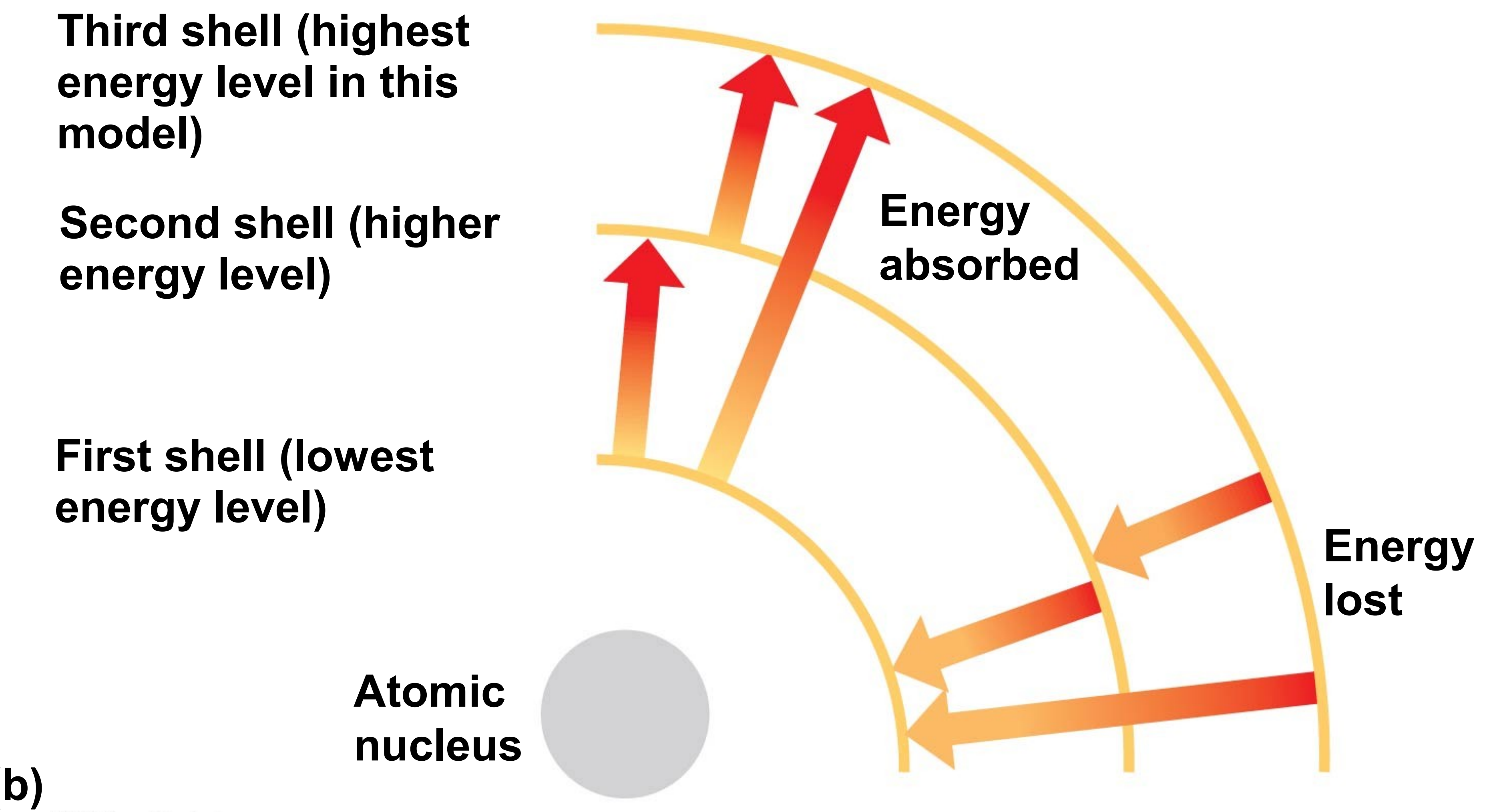
Figure 2.5a

(a) A ball bouncing down a flight of stairs provides an analogy for energy levels of electrons.



- Electrons are found in different **electron shells,** each with a characteristic average distance from the nucleus
- The energy level of each shell increases with distance from the nucleus
- Electrons can move to higher or lower shells by absorbing or releasing energy, respectively

Figure 2.5b



Third shell (highest energy level in this model)
Second shell (higher energy level)
First shell (lowest energy level)
Atomic nucleus
Energy absorbed
Energy lost
(b)

Electron Distribution and Chemical Properties

- The chemical behavior of an atom is determined by the distribution of electrons in electron shells
- The periodic table of the elements shows the electron distribution for each element

Figure 2.6

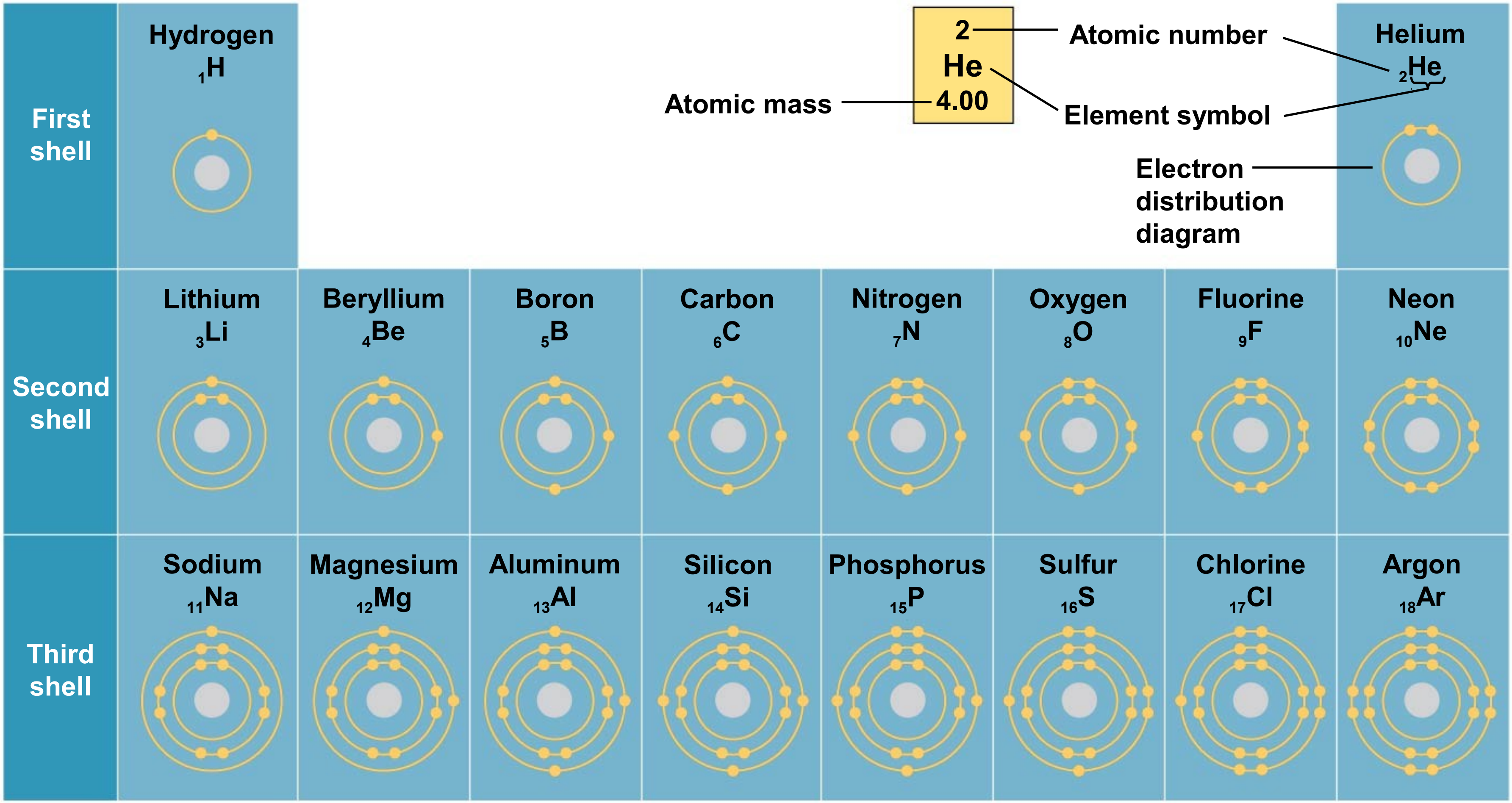
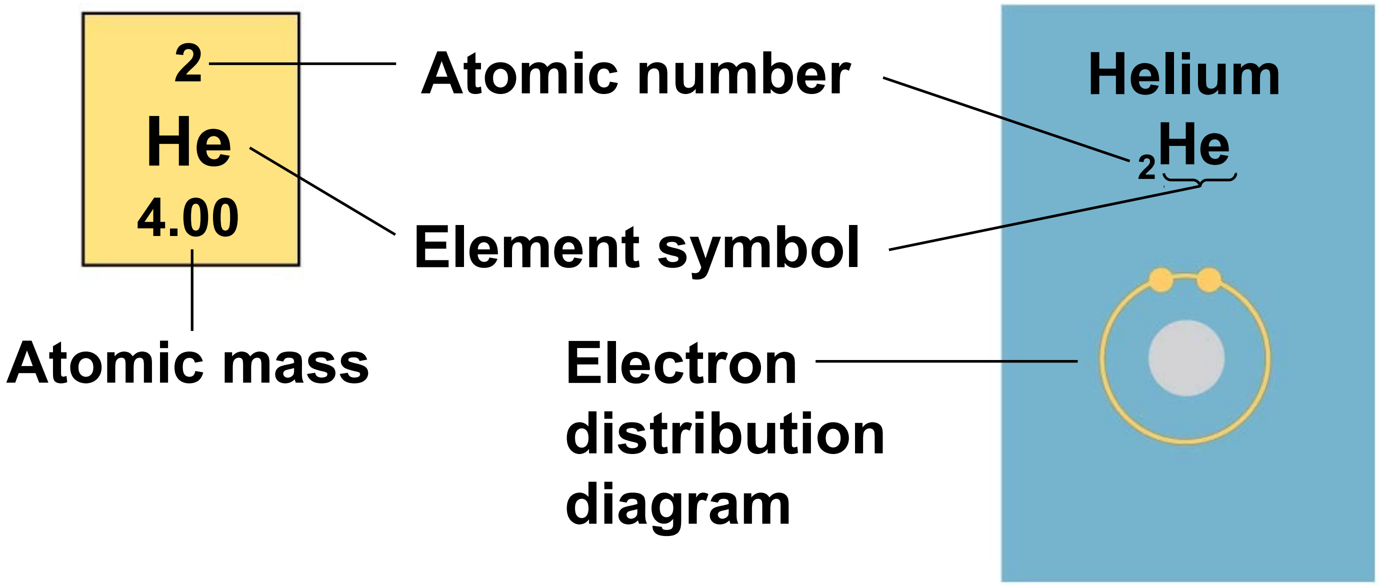

First shell
Second shell
Third shell
Hydrogen
1H
Helium
2He
2
He
4.00
Atomic number
Atomic mass
Element symbol
Electron distribution diagram
Lithium
3Li
Beryllium
4Be
Boron
5B
Carbon
6C
Nitrogen
7N
Oxygen
8O
Fluorine
9F
Neon
10Ne
Sodium
11Na
Magnesium
12Mg
Aluminum
13Al
Silicon
14Si
Phosphorus
15P
Sulfur
16S
Chlorine
17Cl
Argon
18Ar

Figure 2.6a



2
He
4.00
Atomic number
Element symbol
Atomic mass
Helium
$_2$He
Electron distribution diagram

Figure 2.6b

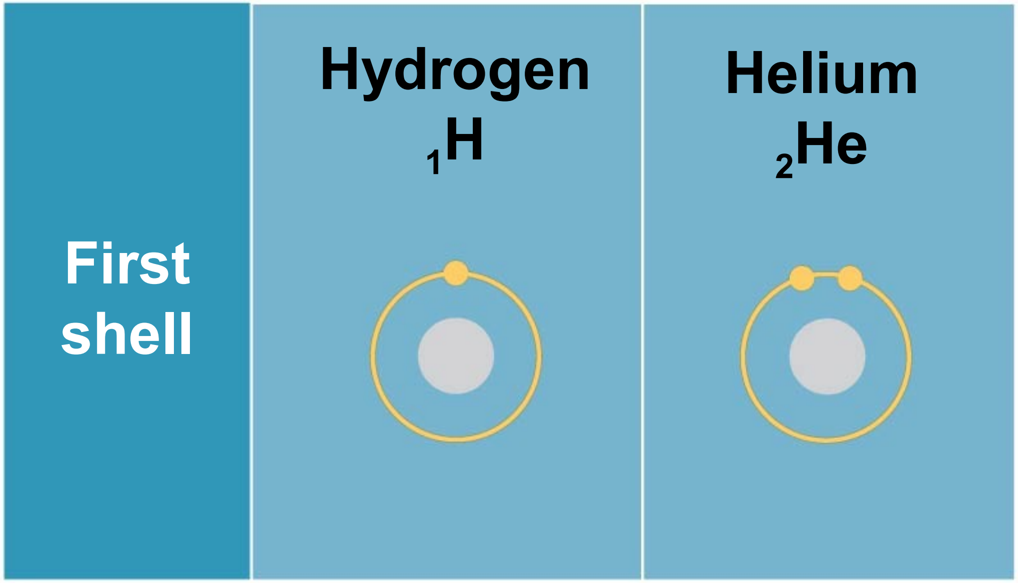
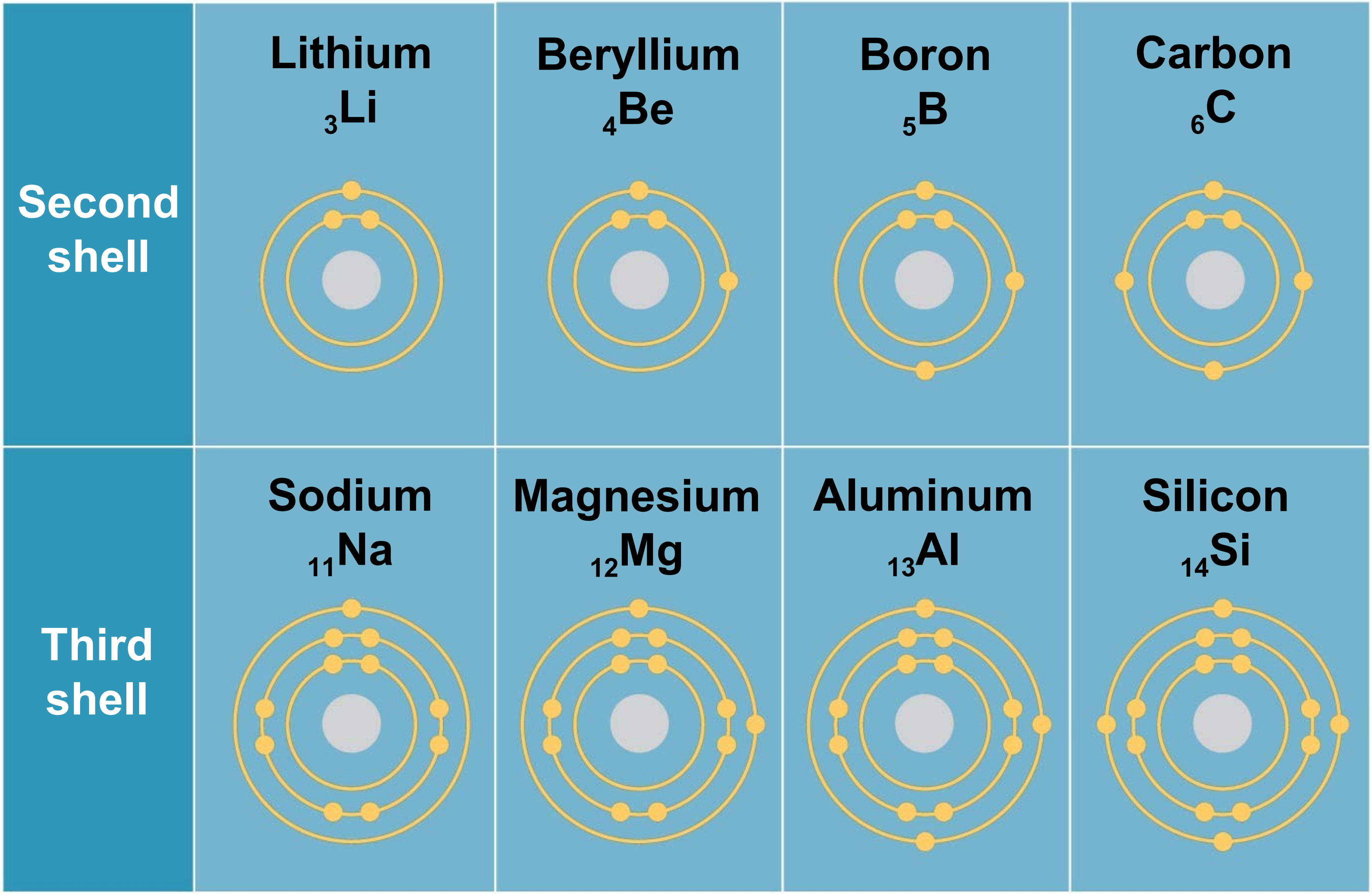
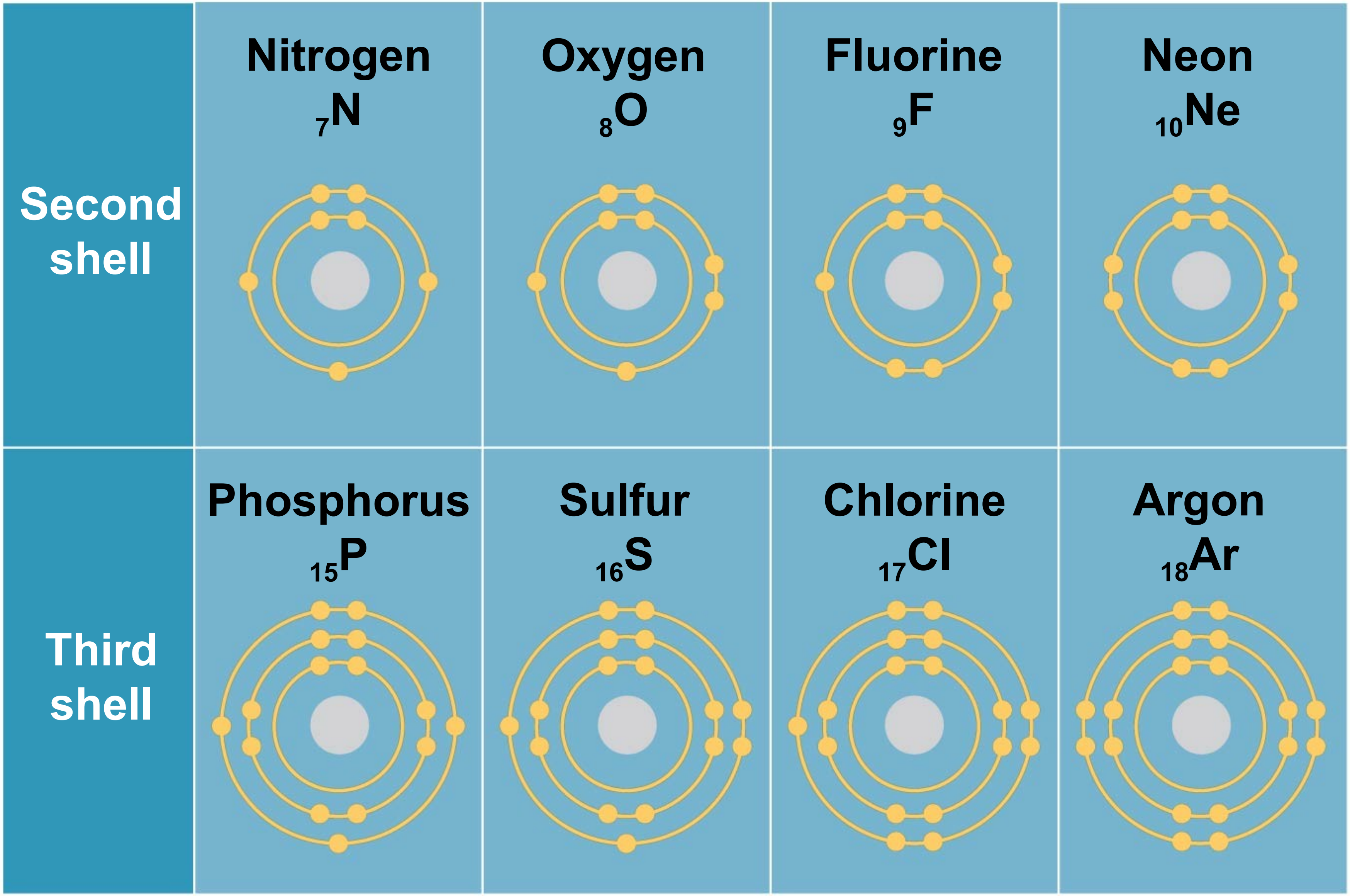


Figure 2.6c



Lithium
$_3$Li
Beryllium
$_4$Be
Boron
$_5$B
Carbon
$_6$C
Second shell
Sodium
$_{11}$Na
Magnesium
$_{12}$Mg
Aluminum
$_{13}$Al
Silicon
$_{14}$Si
Third shell

Figure 2.6d



Second shell
Third shell
Nitrogen
7N
Oxygen
8O
Fluorine
9F
Neon
10Ne
Phosphorus
15P
Sulfur
16S
Chlorine
17Cl
Argon
18Ar

- Chemical behavior of an atom depends mostly on the number of electrons in its outermost shell, or **valence shell**
- **Valence electrons** are those that occupy the valence shell
- The reactivity of an atom arises from the presence of one or more unpaired electrons in the valence shell
- Atoms with completed valence shells are unreactive, or inert

Concept 2.3: The formation and function of molecules depend on chemical bonding between atoms

- Atoms with incomplete valence shells can share or transfer valence electrons with certain other atoms
- This usually results in atoms staying close together, held by attractions called **chemical bonds**

Covalent Bonds

- A **covalent bond** is the sharing of a pair of valence electrons by two atoms
- In a covalent bond, the shared electrons count as part of each atom’s valence shell
- Two or more atoms held together by valence bonds constitute a **molecule**

Figure 2.7-1

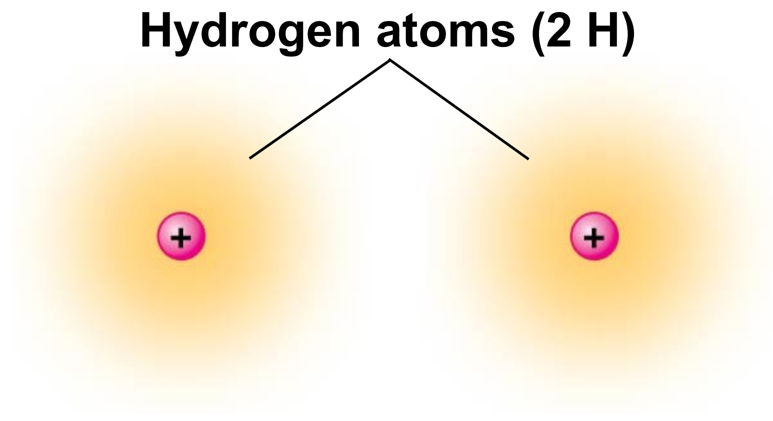


Figure 2.7-2

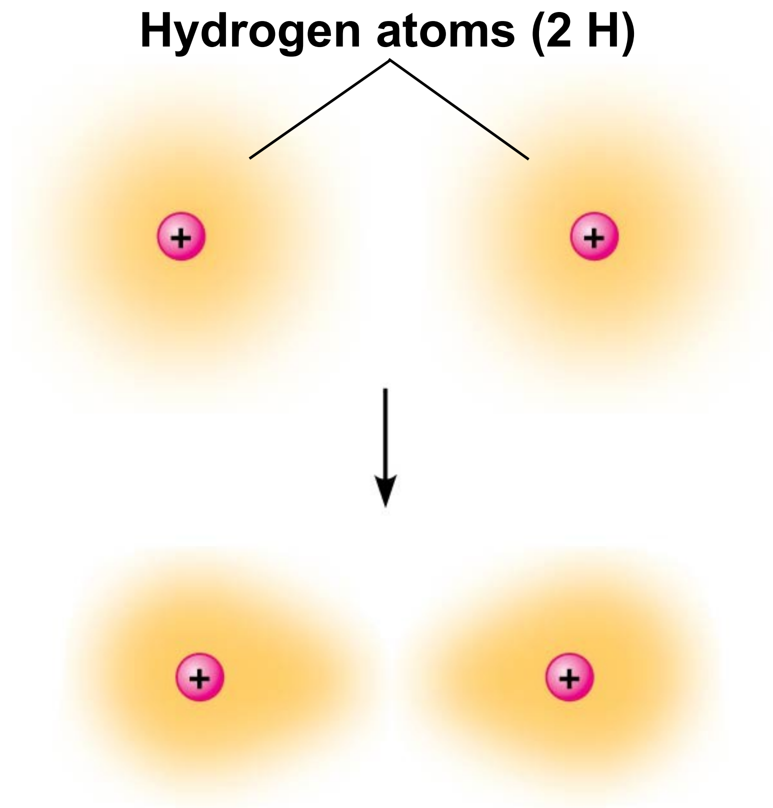
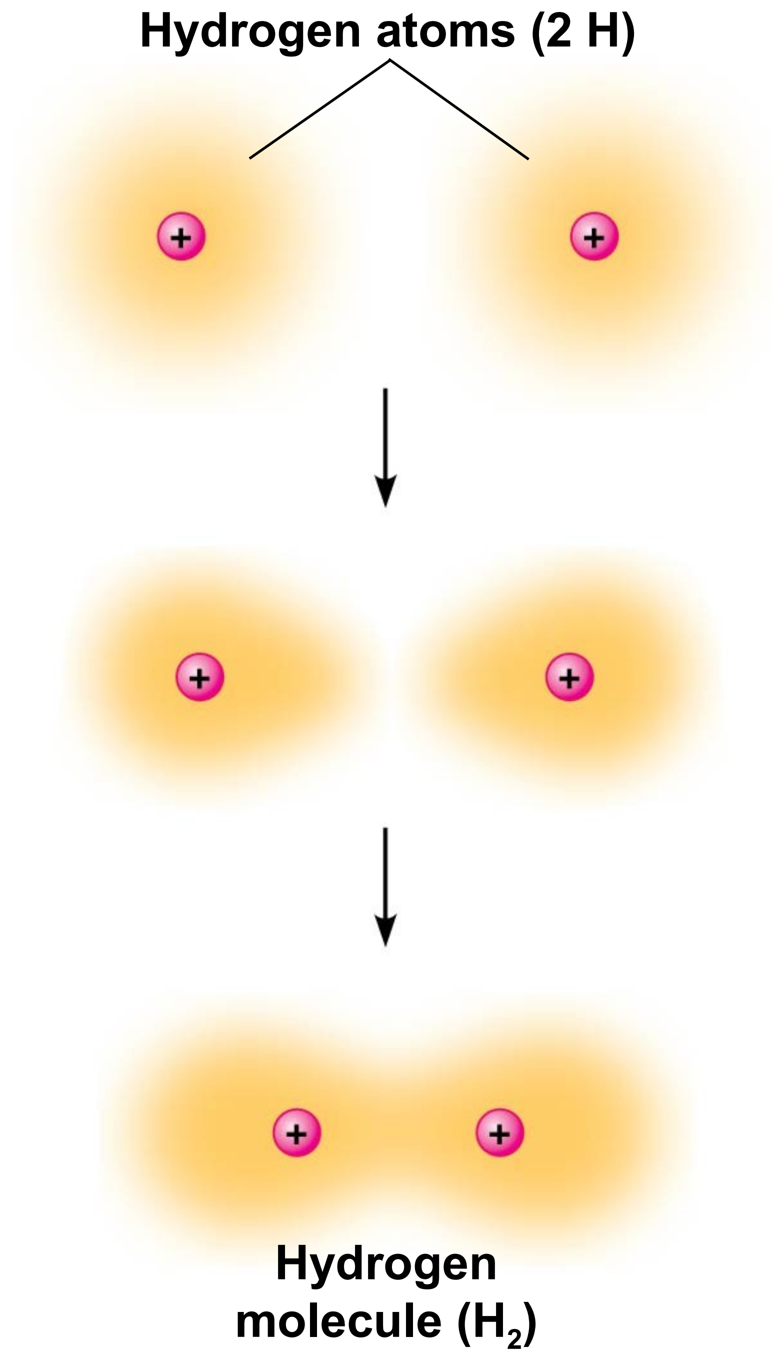


Figure 2.7-3



- The notation used to represent atoms and bonding is called a structural formula
  - For example, H—H
- This can be abbreviated further with a molecular formula
  - For example, H_2

- In a structural formula, a single bond, the sharing of one pair of electrons, is indicated by a single line between the atoms
  - For example, H—H
- A double bond, the sharing of two pairs of electrons, is indicated by a double line between atoms
  - For example, O═O

Figure 2.8

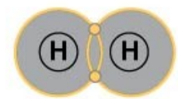
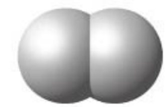
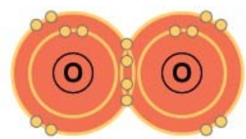
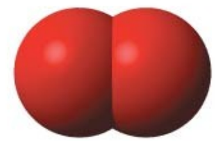
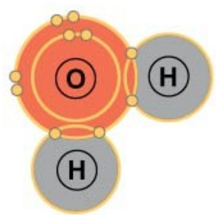
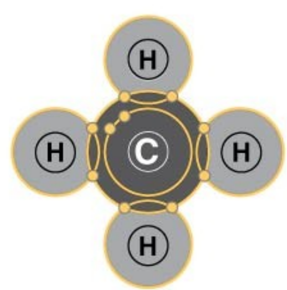
Name and Molecular Formula	Electron Distribution Diagram	Structural Formula	Space-Filling Model
(a) Hydrogen (H_2)		H—H	
(b) Oxygen (O_2)		O=O	
(c) Water (H_2O)		O—H | H	
(d) Methane (CH_4)		H | H—C—H | H	

Figure 2.8a

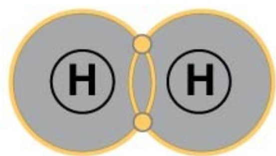
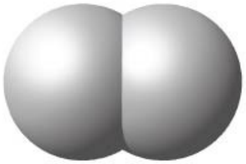
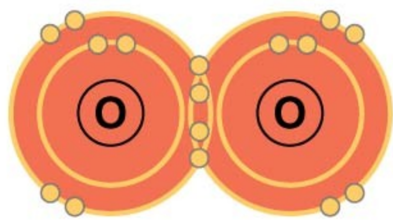
Name and Molecular Formula	Electron Distribution Diagram	Structural Formula	Space-Filling Model
(a) Hydrogen (H_2)		H—H	

Figure 2.8b

Name and Molecular Formula	Electron Distribution Diagram	Structural Formula	Space-Filling Model
(b) Oxygen (O_2)		$O=O$	

- Each atom that can share valence electrons has a bonding capacity, the number of bonds that the atom can form
- Bonding capacity, or **valence**, usually corresponds to the number of electrons required to complete the atom

- Pure elements are composed of molecules of one type of atom, such as H_2 and O_2
- Molecules composed of a combination of two or more types of atoms are called compounds, such as H_2O or CH_4

Figure 2.8c

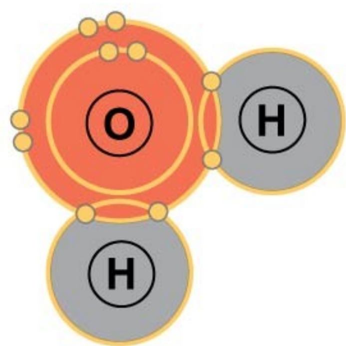
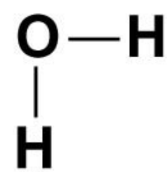
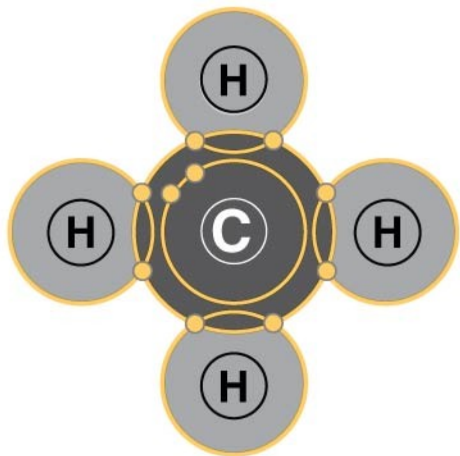
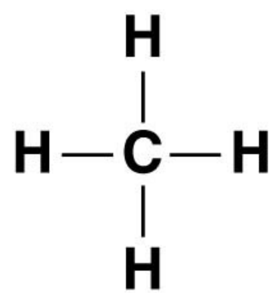
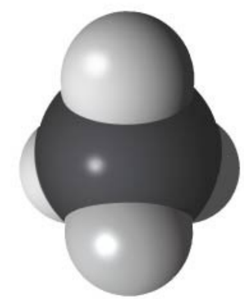
Name and Molecular Formula	Electron Distribution Diagram	Structural Formula	Space-Filling Model
(c) Water (H_2O)			

Figure 2.8d

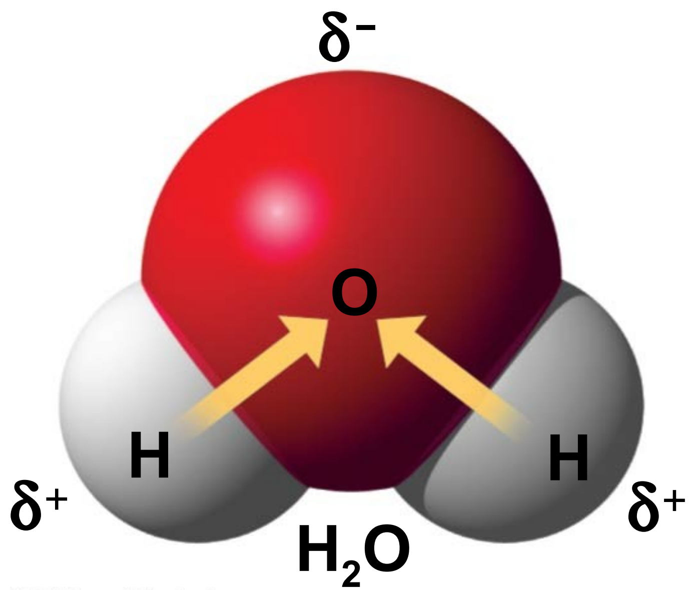
Name and Molecular Formula	Electron Distribution Diagram	Structural Formula	Space-Filling Model
(d) Methane (CH_4)			

- Atoms in a molecule attract electrons to varying degrees
- **Electronegativity** is an atom's attraction for the electrons in a covalent bond
- The more electronegative an atom, the more strongly it pulls shared electrons toward itself

- In a **nonpolar covalent bond**, the atoms share the electron equally
- In a **polar covalent bond**, one atom is more electronegative, and the atoms do not share the electron equally
- Unequal sharing of electrons causes a partial positive or negative charge for each atom or molecule

PLAY Animation: Covalent Bonds

Figure 2.9



Ionic Bonds

- Atoms sometimes strip electrons from their bonding partners
- An example is the transfer of an electron from sodium to chlorine
- After the transfer of an electron, both atoms have charges
- Both atoms also have complete valence shells

Figure 2.10-1

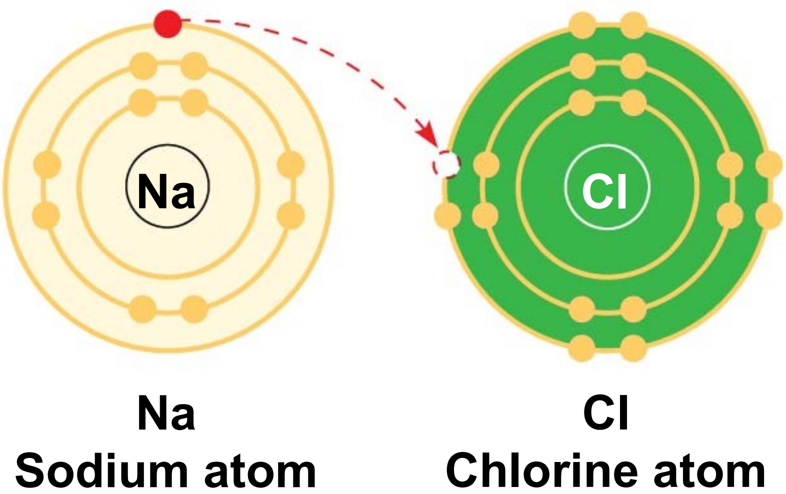
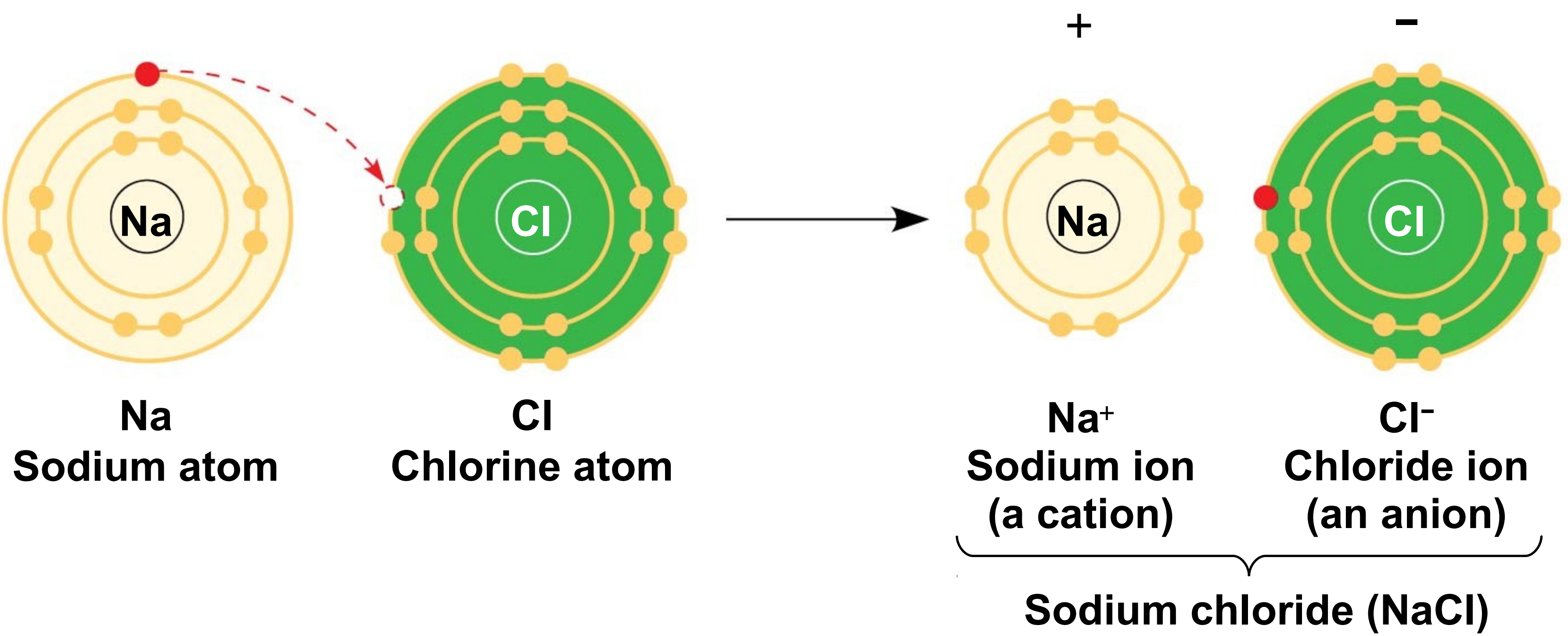


Figure 2.10-2



Na
Cl
+
–
Na
Cl
Na
Sodium atom
Cl
Chlorine atom
Na^+
Sodium ion
(a cation)
Cl^-
Chloride ion
(an anion)
Sodium chloride (NaCl)

- A **cation** is a positively charged ion
- An **anion** is a negatively charged ion
- An **ionic bond** is an attraction between an anion and a cation

- Compounds formed by ionic bonds are called **ionic compounds**, or **salts**
- Salts, such as sodium chloride (table salt), are often found in nature as crystals

PLAY Animation: Ionic Bonds

Figure 2.11

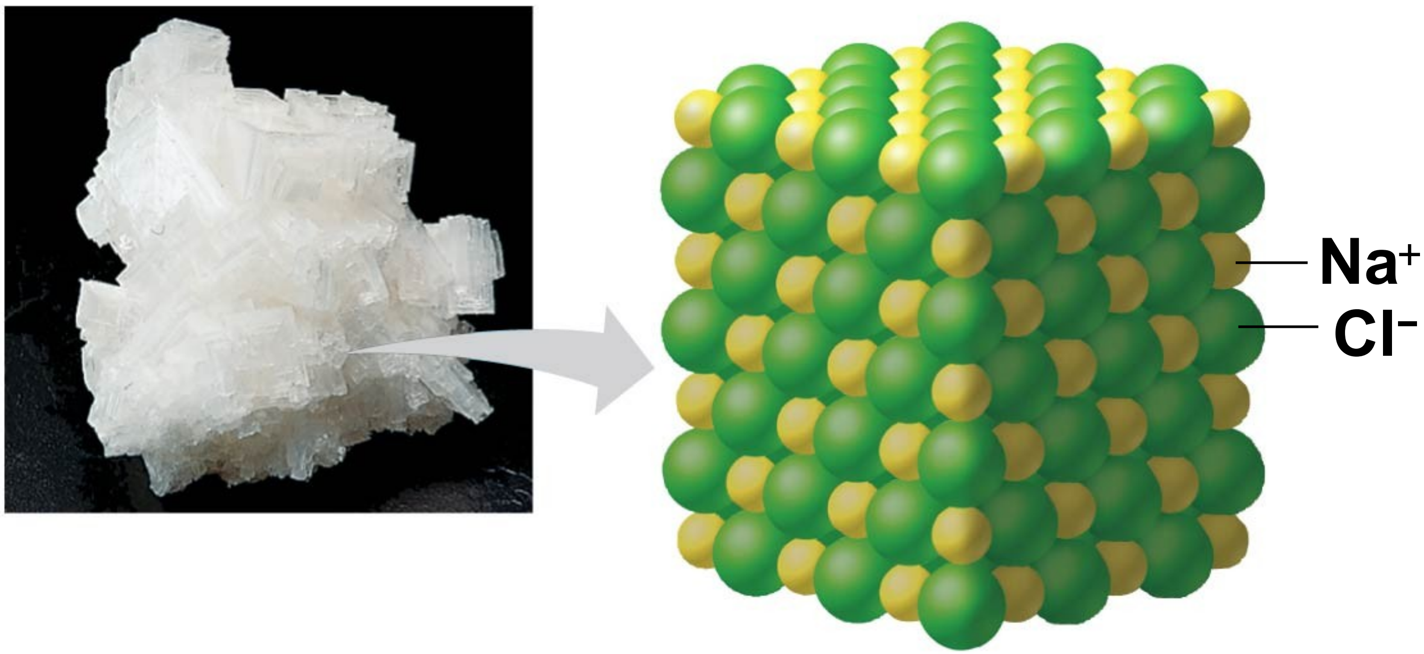


Figure 2.11a

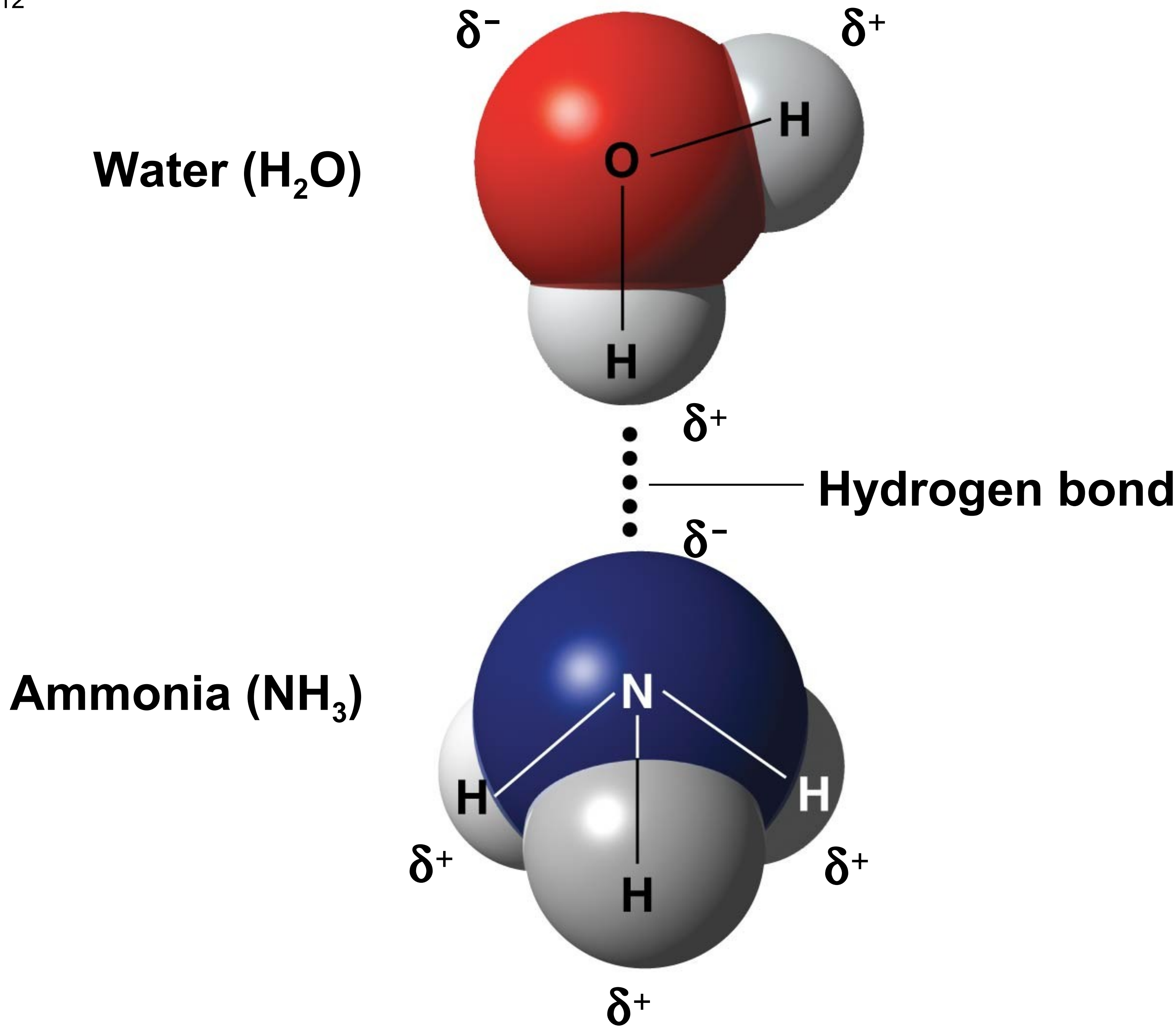
Weak Chemical Bonds

- Most of the strongest bonds in organisms are covalent bonds that form a cell's molecules
- Weak chemical bonds, such as ionic bonds and hydrogen bonds, are also important
- Many large biological molecules are held in their functional form by weak bonds

Hydrogen Bonds

- A **hydrogen bond** forms when a hydrogen atom covalently bonded to one electronegative atom is also attracted to another electronegative atom
- In living cells, the electronegative partners are usually oxygen or nitrogen atoms

Figure 2.12

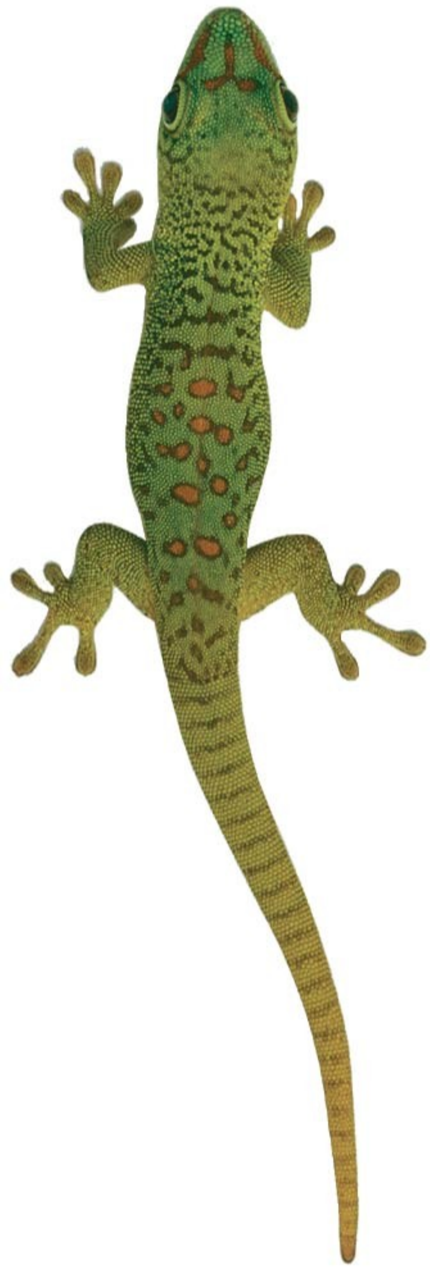


Van der Waals Interactions

- If electrons are distributed asymmetrically in molecules or atoms, they can result in “hot spots” of positive or negative charge
- **Van der Waals interactions** are attractions between molecules that are close together as a result of these charges

- Van der Waals interactions are individually weak and occur only when atoms and molecules are very close together
- Collectively, such interactions can be strong, as between molecules of a gecko’s toe hairs and a wall surface

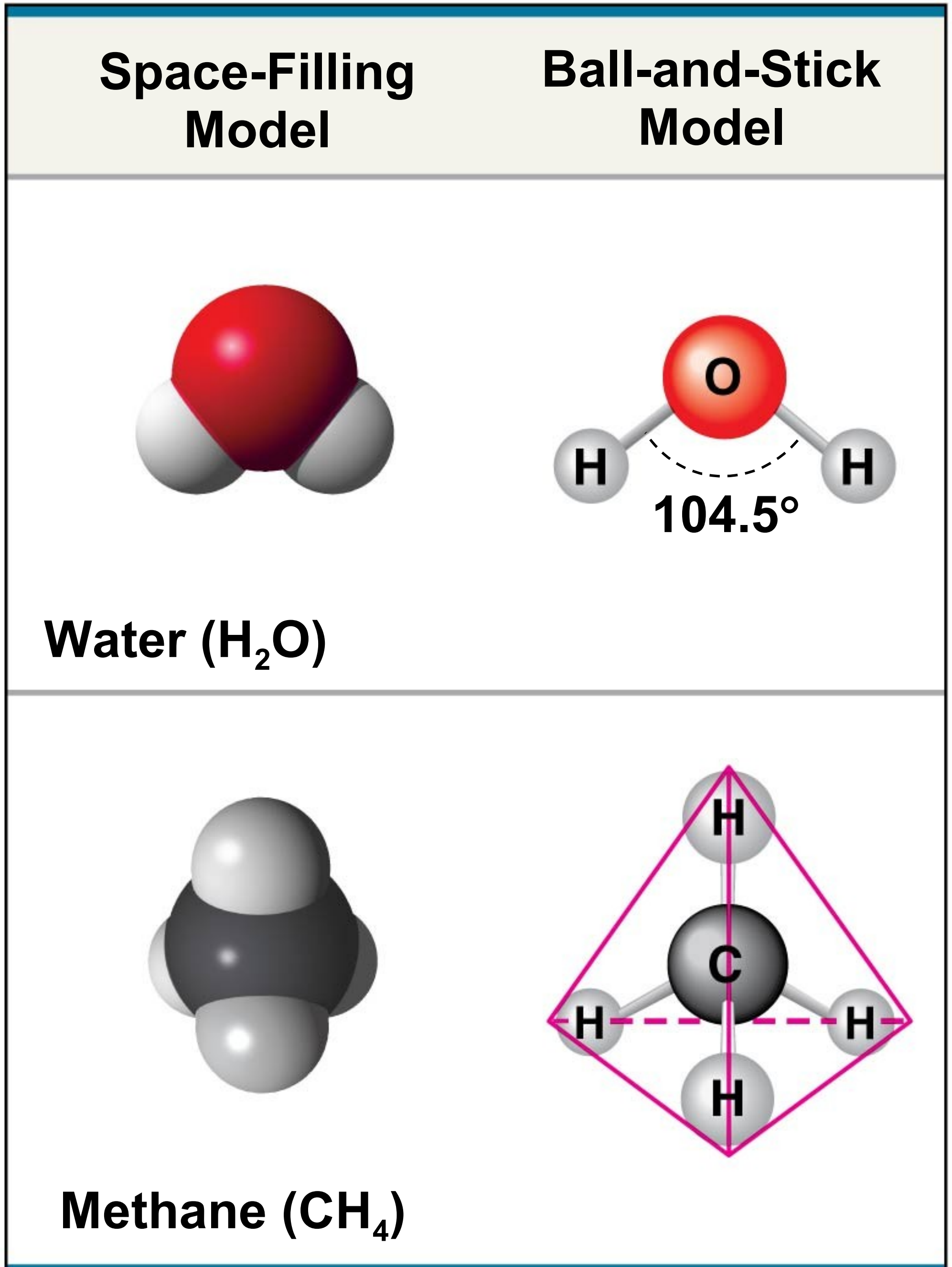
Figure 2.UN01



Molecular Shape and Function

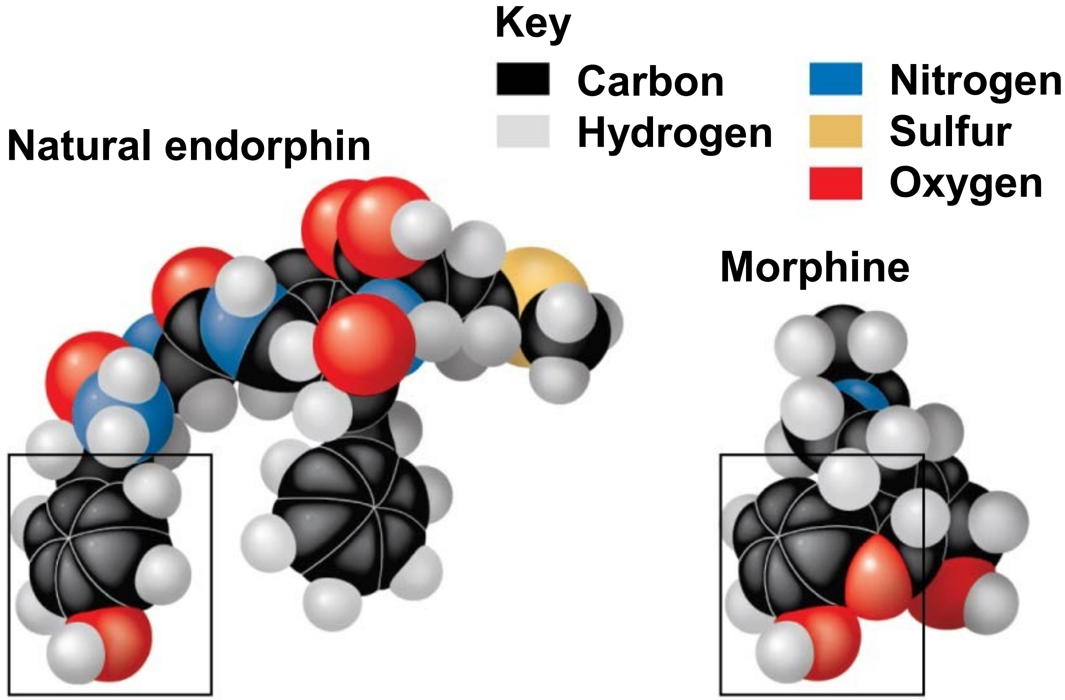
- A molecule's shape is usually very important to its function
- Molecular shape determines how biological molecules recognize and respond to one another

Figure 2.13

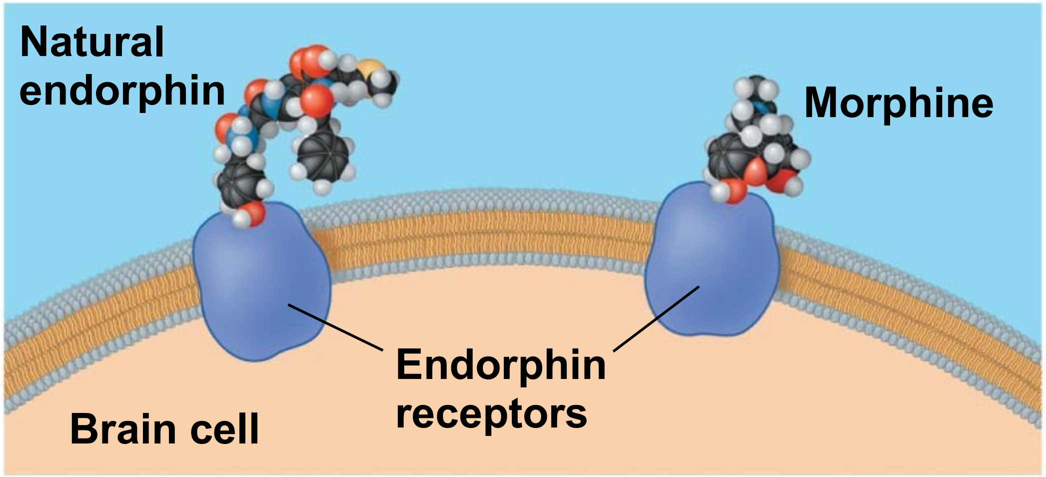

Space-Filling Model
Ball-and-Stick Model
O
H
H
104.5°
Water (H_2O)
H
C
H
H
H
Methane (CH_4)

- Biological molecules recognize and interact with each other with a specificity based on molecular shape
- Molecules with similar shapes can have similar biological effects

Figure 2.14

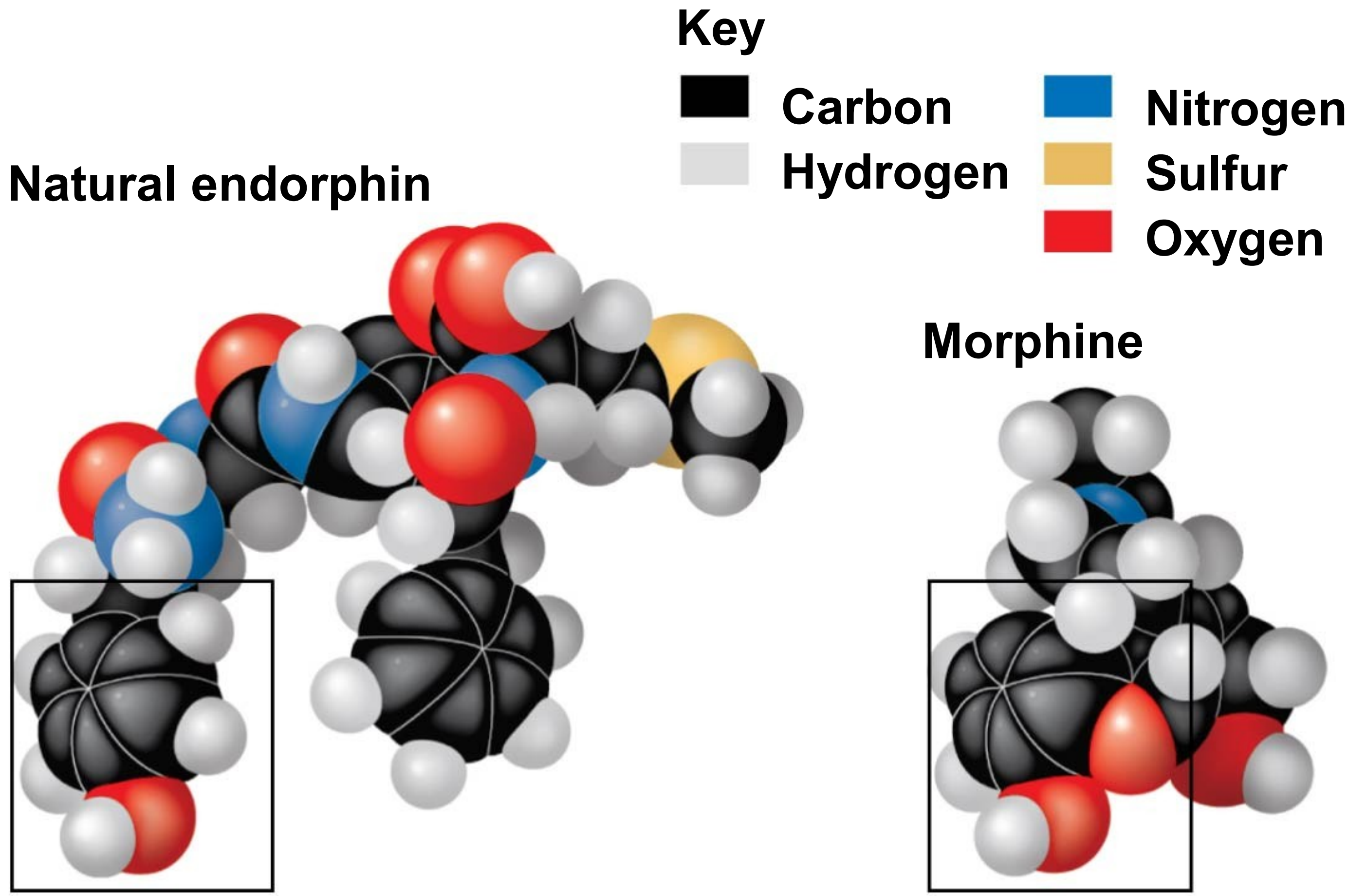


(a) Structures of endorphin and morphine



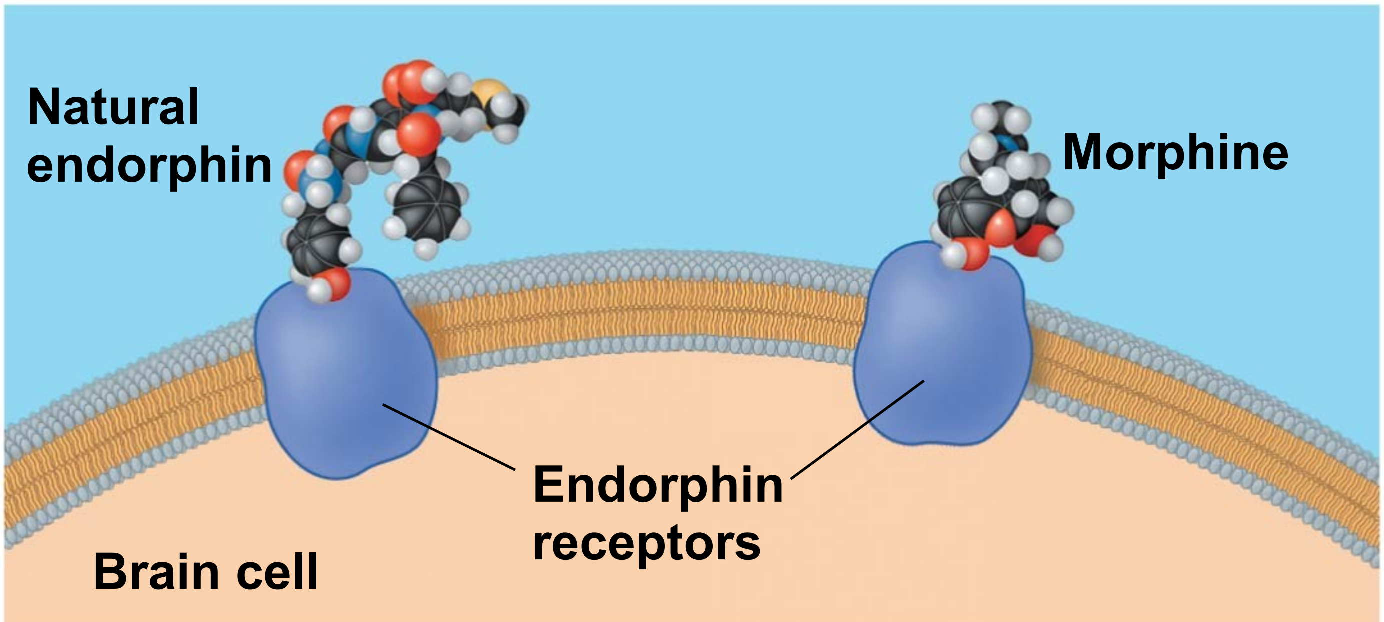
(b) Binding to endorphin receptors

Figure 2.14a



(a) Structures of endorphin and morphine

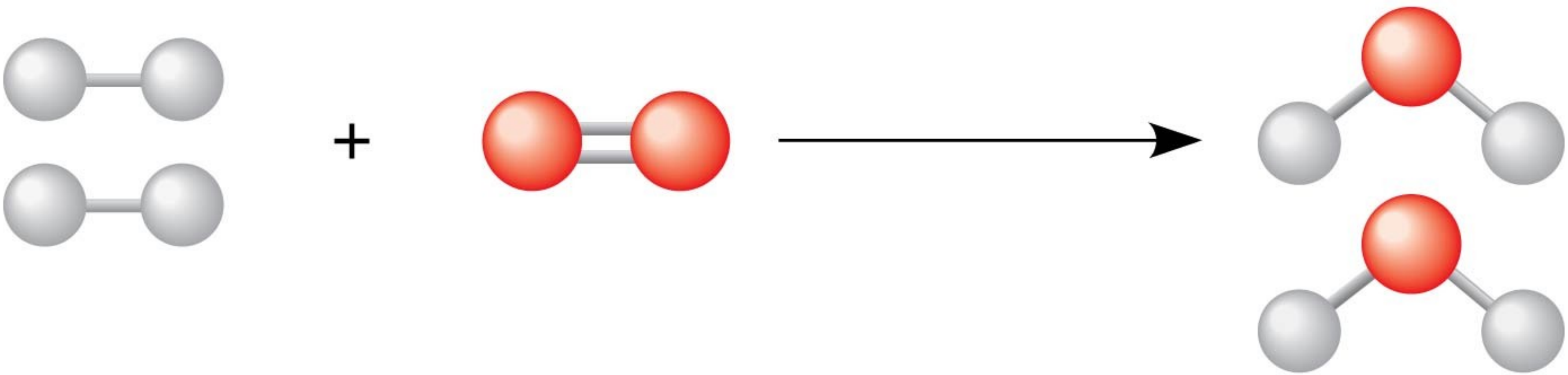
Figure 2.14b



(b) Binding to endorphin receptors

Concept 2.4: Chemical reactions make and break chemical bonds

- **Chemical reactions** are the making and breaking of chemical bonds
- The starting molecules of a chemical reaction are called **reactants**
- The final molecules of a chemical reaction are called **products**



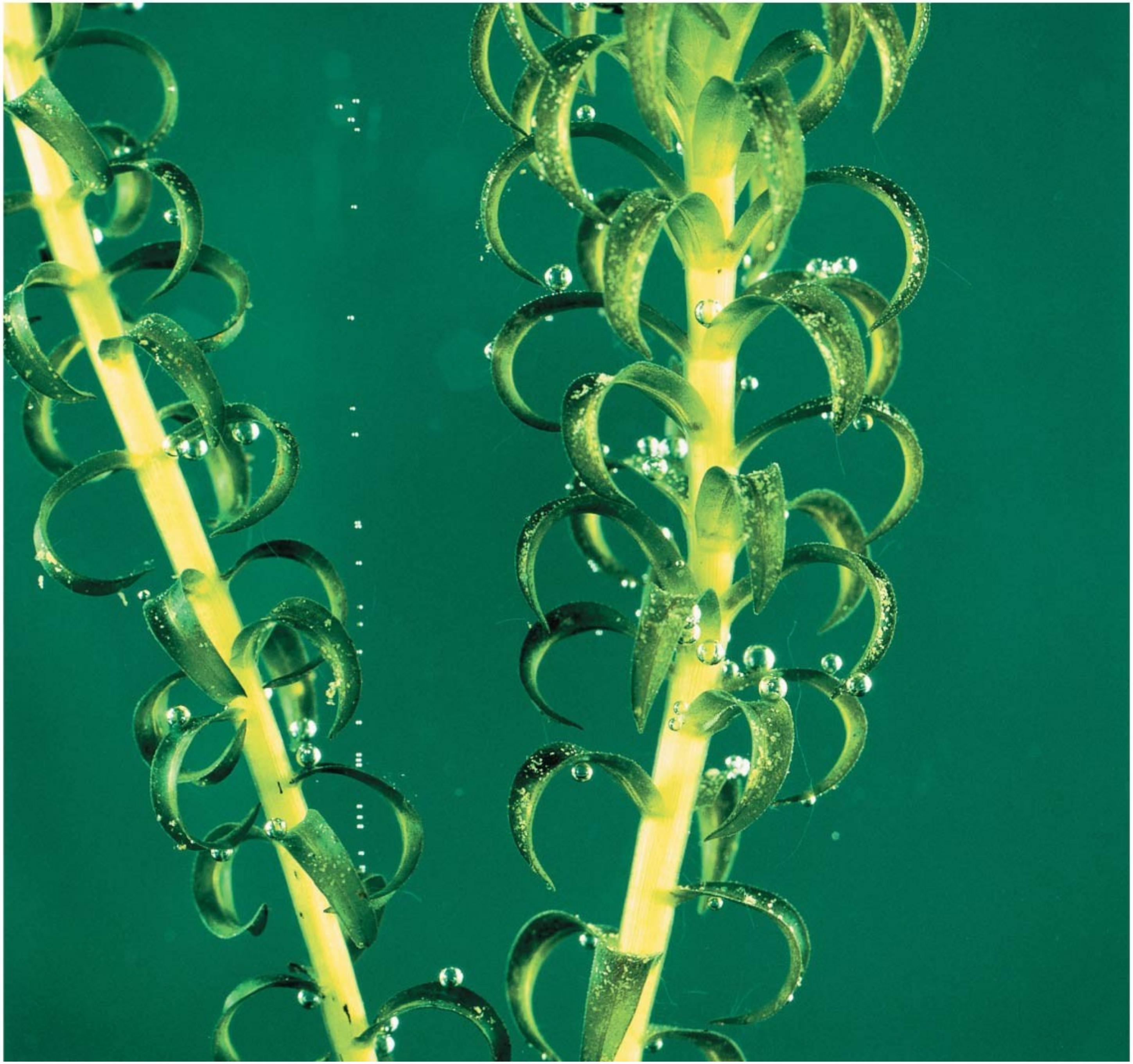
$$2\,H_2 + O_2 \rightarrow 2\,H_2O$$

Reactants **Reaction** **Products**

- Photosynthesis is an important chemical reaction
- Sunlight powers the conversion of carbon dioxide and water to glucose and oxygen

$$6\ CO_2 + 6\ H_2O \rightarrow C_6H_{12}O_6 + 6\ O_2$$

Figure 2.15

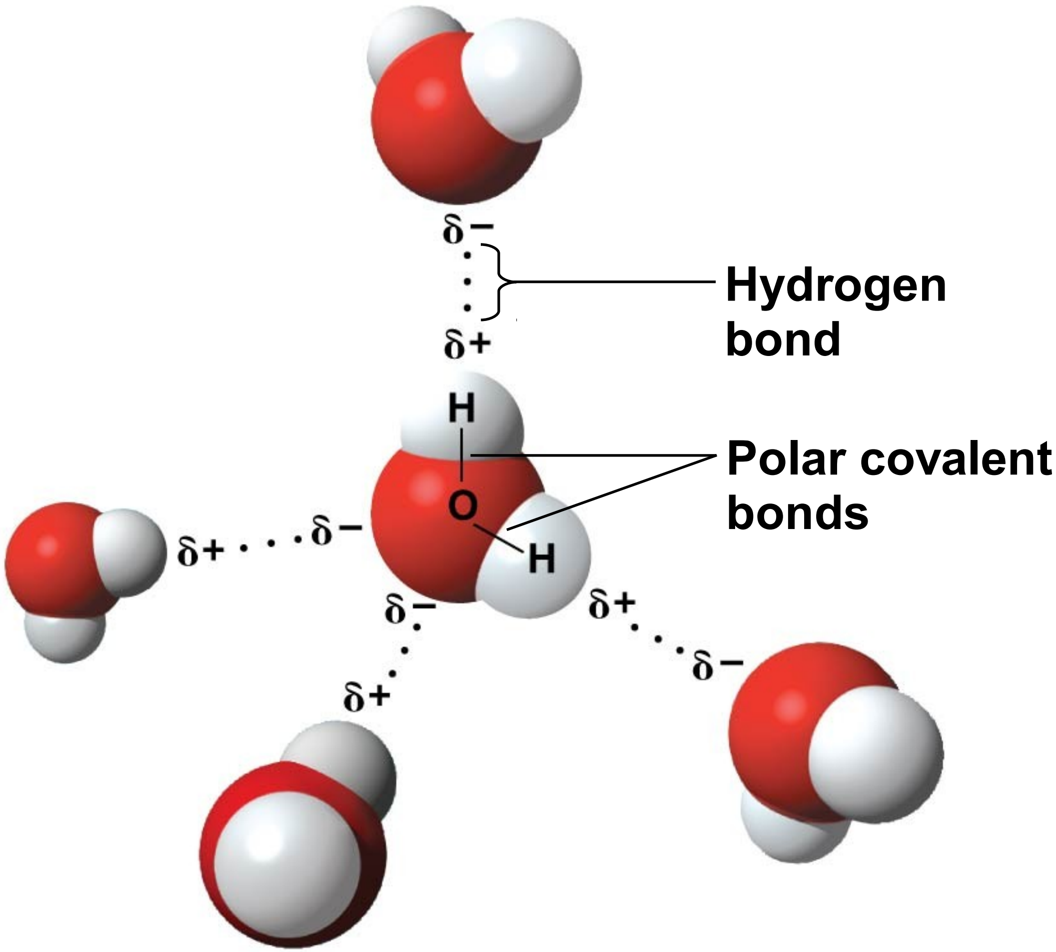


- All chemical reactions are reversible: Products of the forward reaction become reactants for the reverse reaction
- **Chemical equilibrium** is reached when the forward and reverse reaction rates are equal

Concept 2.5: Hydrogen bonding gives water properties that help make life possible on Earth

- All organisms are made mostly of water and live in an environment dominated by water
- Water molecules are polar, with the oxygen region having a partial negative charge ($\delta-$) and the hydrogen region a slight positive charge ($\delta+$)
- Two water molecules are held together by a hydrogen bond

Figure 2.16



- Four emergent properties of water contribute to Earth’s suitability for life:
  - Cohesive behavior
  - Ability to moderate temperature
  - Expansion upon freezing
  - Versatility as a solvent

Cohesion of Water Molecules

- Water molecules are linked by multiple hydrogen bonds
- The molecules stay close together because of this; it is called **cohesion**

- Cohesion due to hydrogen bonding contributes to the transport of water and nutrients against gravity in plants
- **Adhesion**, the clinging of one substance to another, also plays a role

PLAY Animation: Water Structure

Figure 2.17

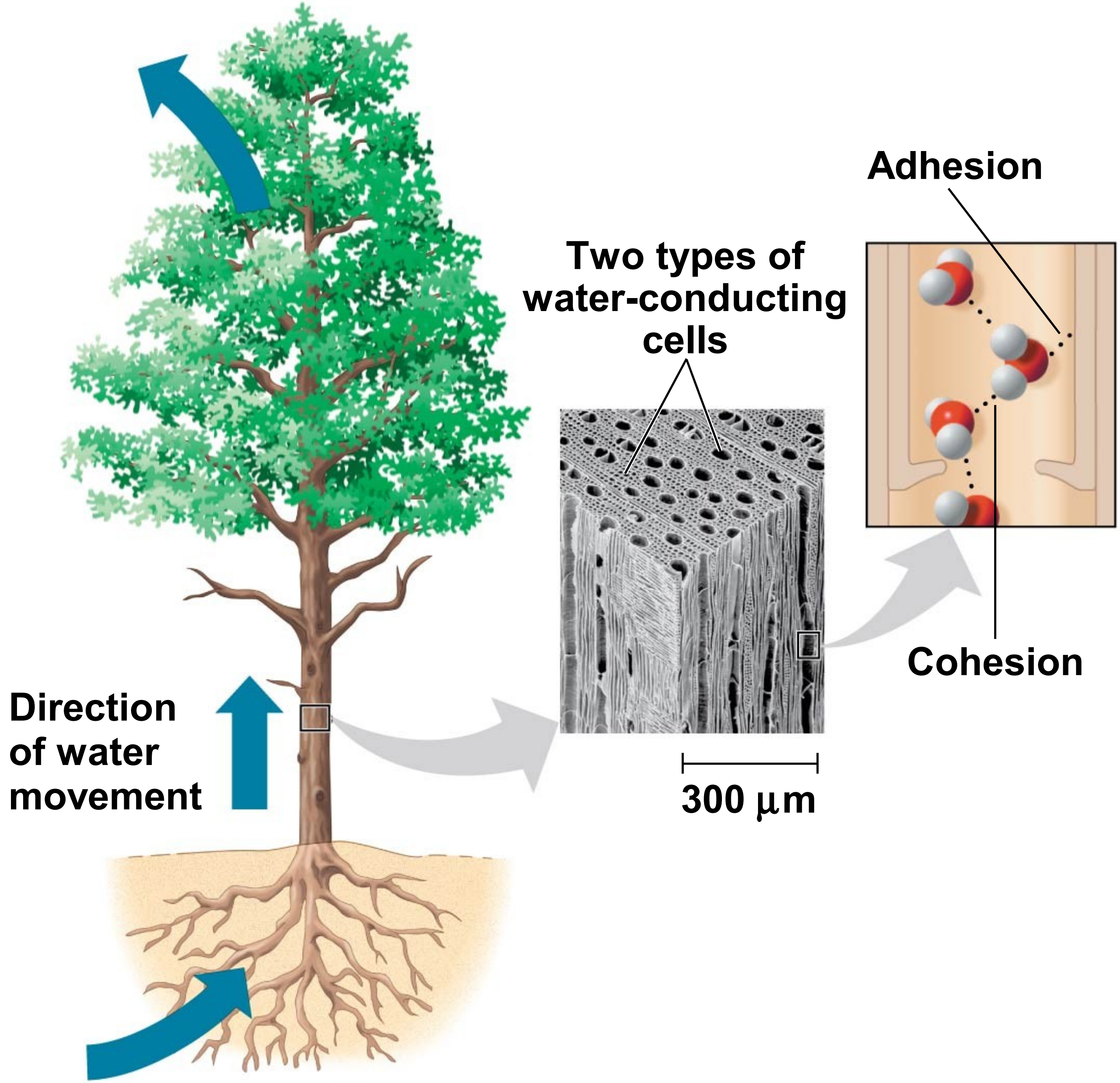
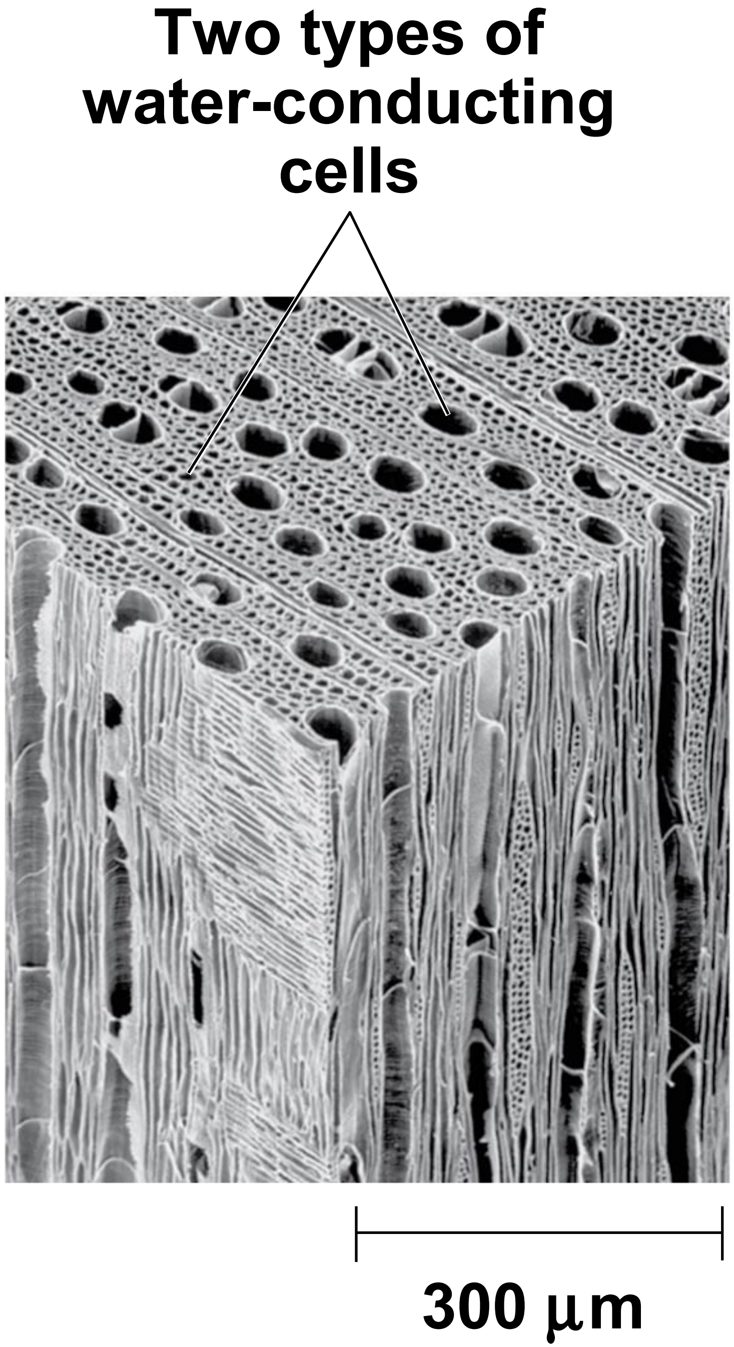

Direction
of water
movement
Two types of
water-conducting
cells
300 µm
Adhesion
Cohesion

Figure 2.17a



- **Surface tension** is a measure of how hard it is to break the surface of a liquid
- Surface tension is related to cohesion

PLAY Animation: Water Transport

PLAY Animation: Water Transport in Plants

Figure 2.18

Moderation of Temperature by Water

- Water absorbs heat from warmer air and releases stored heat to cooler air
- Water can absorb or release a large amount of heat with only a slight change in its own temperature

Temperature and Heat

- **Kinetic energy** is the energy of motion
- **Thermal energy** is a measure of the total amount of kinetic energy due to molecular motion
- **Temperature** represents the average kinetic energy of molecules
- Thermal energy in transfer from one body of matter to another is defined as **heat**

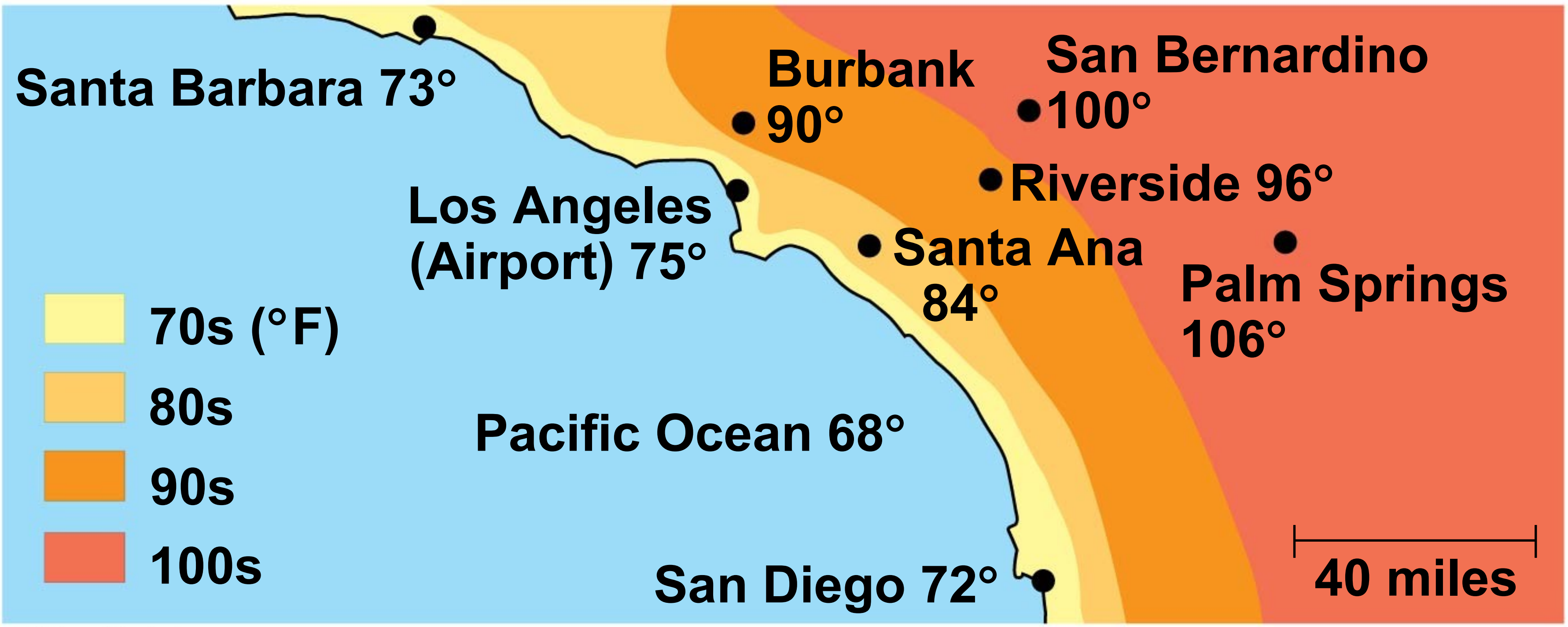
- The **Celsius scale** is a measure of temperature using Celsius degrees (°C)
- A **calorie (cal)** is the amount of heat required to raise the temperature of 1 g of water by 1°C
- The "calories" on food packages are actually **kilocalories (kcal)**, where 1 kcal = 1,000 cal
- The **joule (J)** is another unit of energy, where 1 J = 0.239 cal, or 1 cal = 4.184 J

Water's High Specific Heat

- The **specific heat** of a substance is the amount of heat that must be absorbed or lost for 1 g of that substance to change its temperature by 1°C
- The specific heat of water is 1 cal/g/°C
- Water resists changing its temperature because of its high specific heat

- Water's high specific heat can be traced to hydrogen bonding
  - Heat is absorbed when hydrogen bonds break
  - Heat is released when hydrogen bonds form
- The high specific heat of water keeps temperature fluctuations within limits that permit life

Figure 2.19


Santa Barbara 73°
Burbank
90°
San Bernardino
100°
Riverside 96°
Los Angeles
(Airport) 75°
Santa Ana
84°
Palm Springs
106°
70s (°F)
80s
90s
100s
Pacific Ocean 68°
San Diego 72°
40 miles

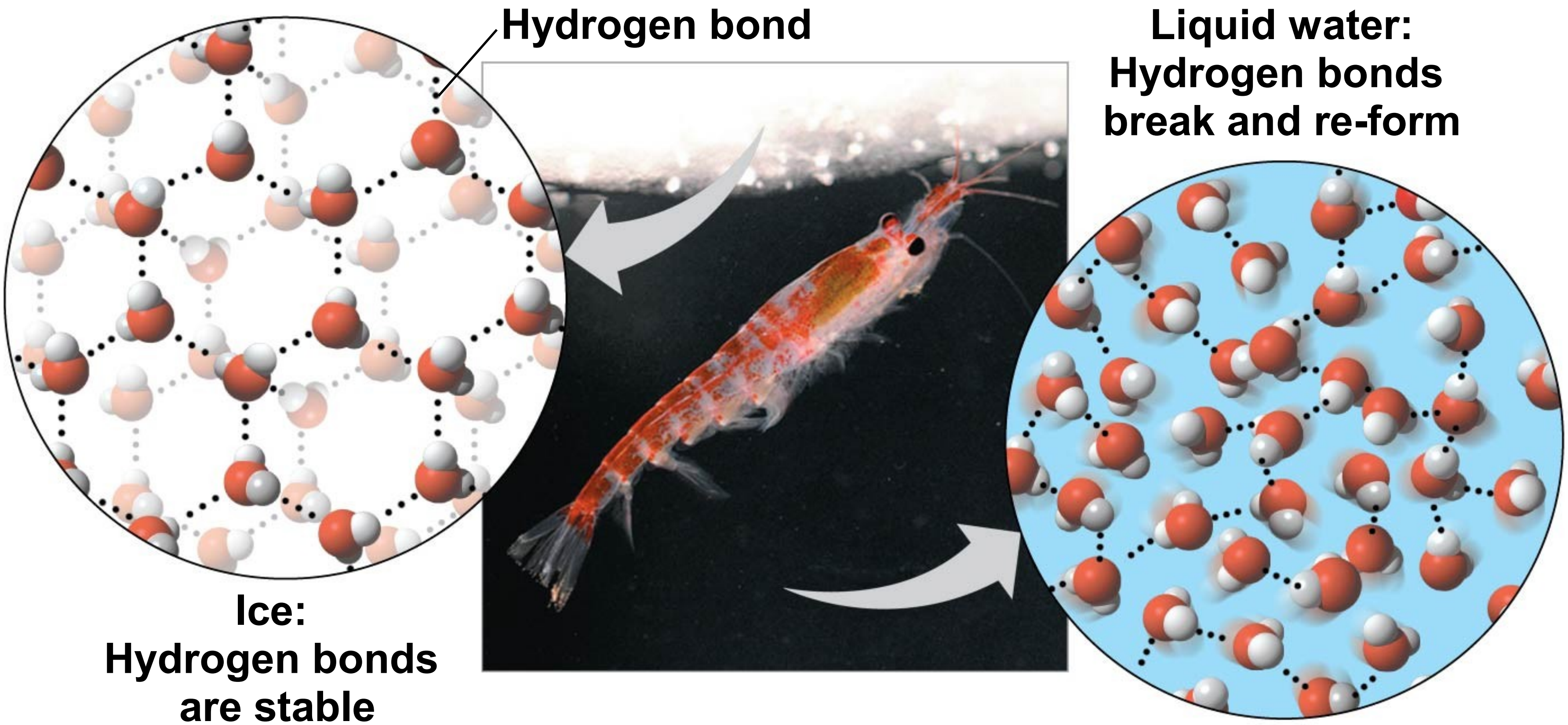
Evaporative Cooling

- Evaporation is transformation of a substance from liquid to gas
- **Heat of vaporization** is the heat a liquid must absorb for 1 g to be converted to gas
- As a liquid evaporates, its remaining surface cools, a process called **evaporative cooling**
- Evaporative cooling of water helps stabilize temperatures in organisms and bodies of water

Floating of Ice on Liquid Water

- Ice floats in liquid water because hydrogen bonds in ice are more "ordered," making ice less dense
- Water reaches its greatest density at 4°C
- If ice sank, all bodies of water would eventually freeze solid, making life impossible on Earth

Figure 2.20



Hydrogen bond
Liquid water:
Hydrogen bonds
break and re-form
Ice:
Hydrogen bonds
are stable

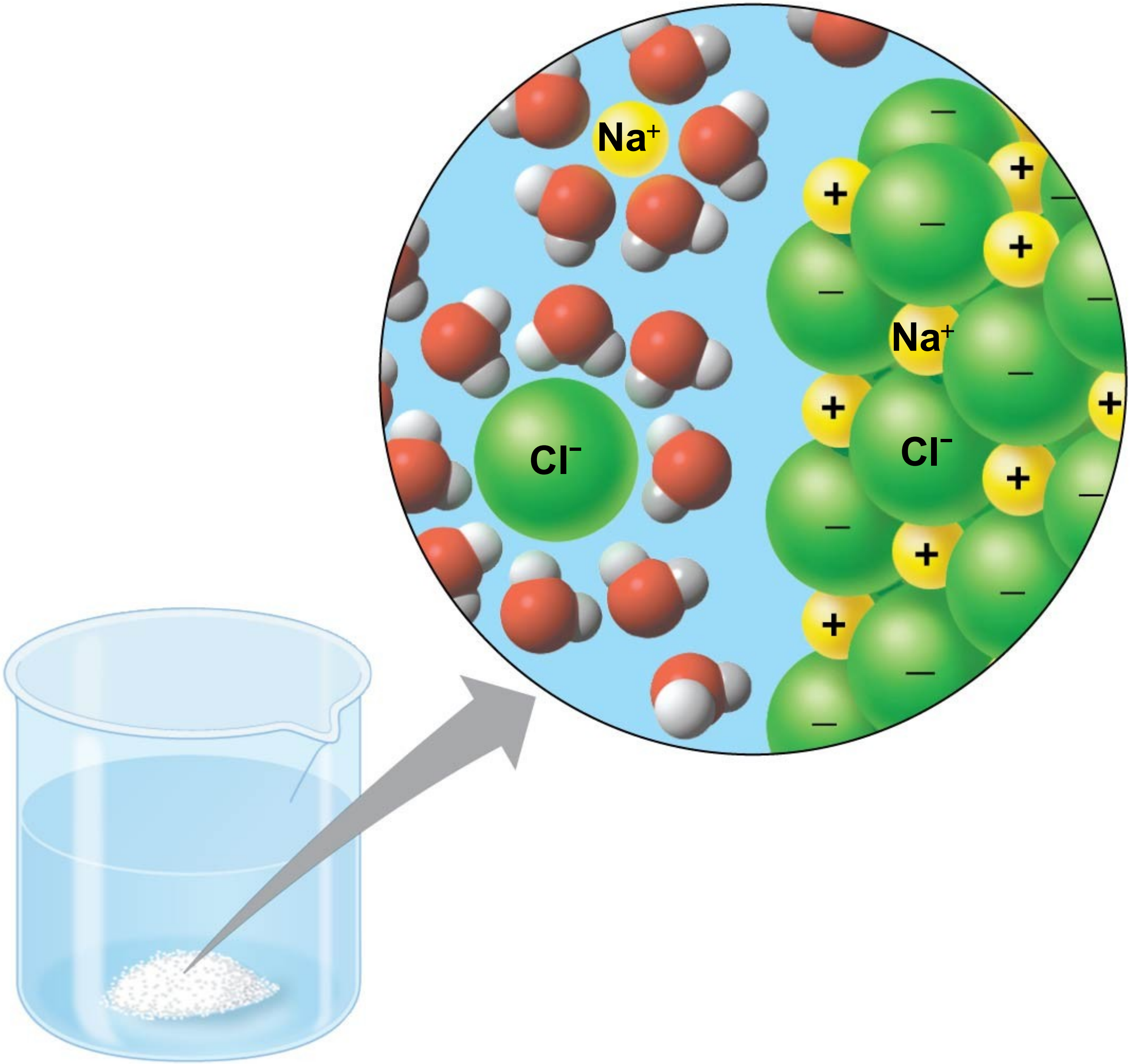
Figure 2.20a

Water: The Solvent of Life

- A **solution** is a liquid that is a homogeneous mixture of substances
- A **solvent** is the dissolving agent of a solution
- The **solute** is the substance that is dissolved
- An **aqueous solution** is one in which water is the solvent

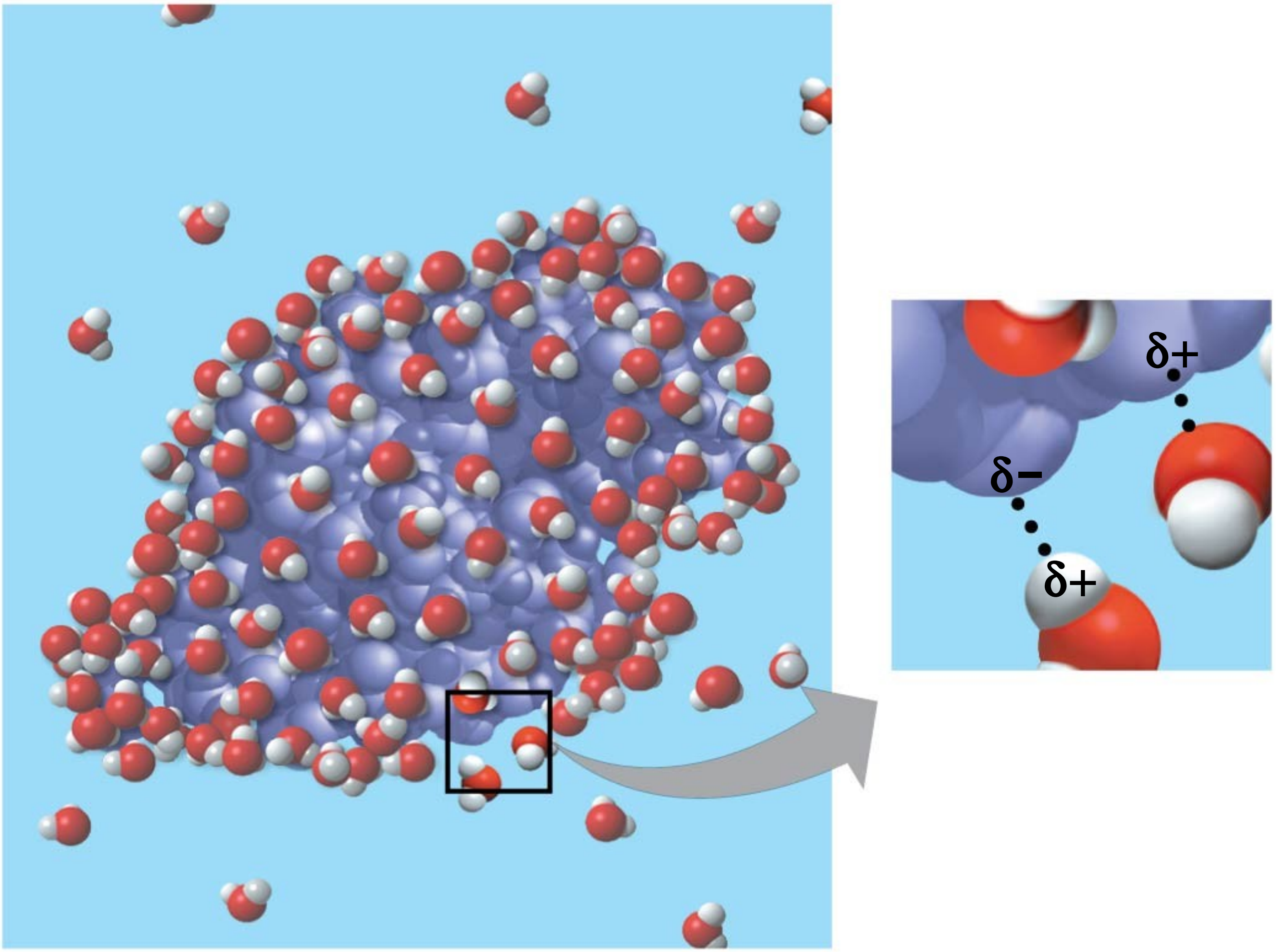
- Water is a versatile solvent due to its polarity, which allows it to form hydrogen bonds easily
- When an ionic compound is dissolved in water, each ion is surrounded by a sphere of water molecules called a **hydration shell**

Figure 2.21



- Water can also dissolve compounds made of nonionic polar molecules
- Even large polar molecules such as proteins can dissolve in water if they have ionic and polar regions

Figure 2.22



Hydrophilic and Hydrophobic Substances

- A **hydrophilic** substance is one that has an affinity for water
- A **hydrophobic** substance is one that does not have an affinity for water
- Oil molecules are hydrophobic because they have relatively nonpolar bonds
- A **colloid** is a stable suspension of fine particles in a liquid

Solute Concentration in Aqueous Solutions

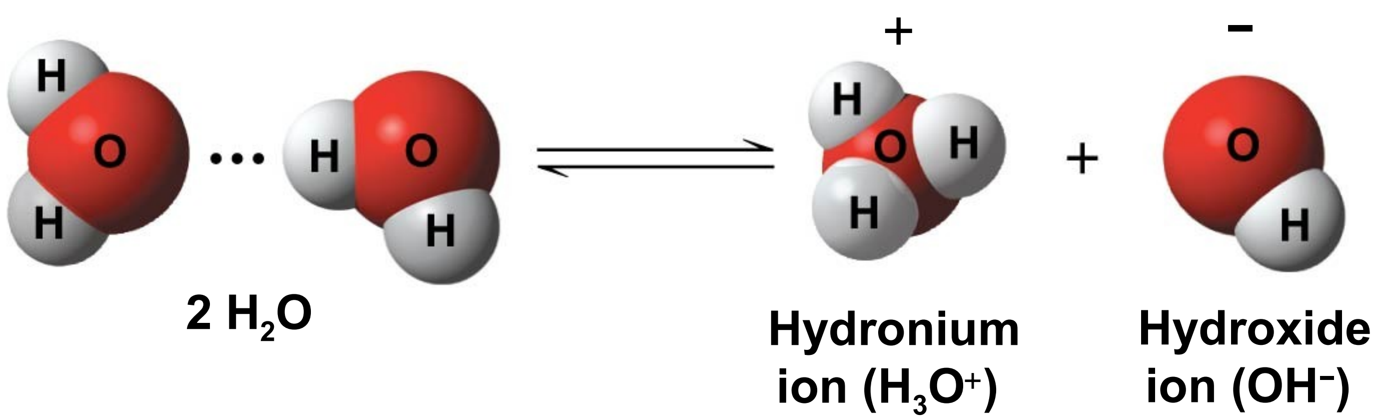
- Most biochemical reactions occur in water
- Chemical reactions depend on collisions of molecules and therefore on the concentration of solutes in an aqueous solution

- **Molecular mass** is the sum of all masses of all atoms in a molecule
- Numbers of molecules are usually measured in moles, where 1 **mole (mol)** $= 6.02 \times 10^{23}$ molecules
- Avogadro's number and the unit *dalton* were defined such that 6.02×10^{23} daltons $= 1$ g
- **Molarity (*M*)** is the number of moles of solute per liter of solution

Acids and Bases

- Sometimes a **hydrogen ion** (H^+) is transferred from one water molecule to another, leaving behind a **hydroxide ion** (OH^-)
- The proton (H^+) binds to the other water molecule, forming **a hydronium ion** (H_3O^+)
- By convention, H^+ is used to represent the hydronium ion

Figure 2.UN03



- Though water dissociation is rare and reversible, it is important in the chemistry of life
- H^+ and OH^- are very reactive
- Solutes called acids and bases disrupt the balance between H^+ and OH^- in pure water
- **Acids** increase the H^+ concentration in water, while **bases** reduce the concentration of H^+

- An **acid** is any substance that increases the H^+ concentration of a solution
- A **base** is any substance that reduces the H^+ concentration of a solution

- A strong acid like hydrochloric acid, HCl, dissociates completely into H^+ and Cl^- in water:

$$HCl \rightarrow H^+ + Cl^-$$

- Sodium hydroxide, NaOH, acts as a strong base indirectly by dissociating completely to form hydroxide ions
- These combine with H^+ ions to form water:

$$NaOH \rightarrow Na^+ + OH^-$$

- Ammonia, NH_3, acts as a relatively weak base when it attracts an H^+ ion from the solution and forms ammonium, NH_4^+
- This is a reversible reaction, as shown by the double arrows:

$$NH_3 + H^+ \rightleftharpoons NH_4^+$$

- Carbonic acid, H_2CO_3, acts as a weak acid, which can reversibly release and accept back H^+ ions:

$$H_2CO_3 \rightleftharpoons HCO_3^- + H^+$$

The pH Scale

- In any aqueous solution at 25°C, the product of H^+ and OH^- is constant and can be written as

$$[H^+][OH^-] = 10^{-14}$$

- The **pH** of a solution is defined by the negative logarithm of H^+ concentration, written as

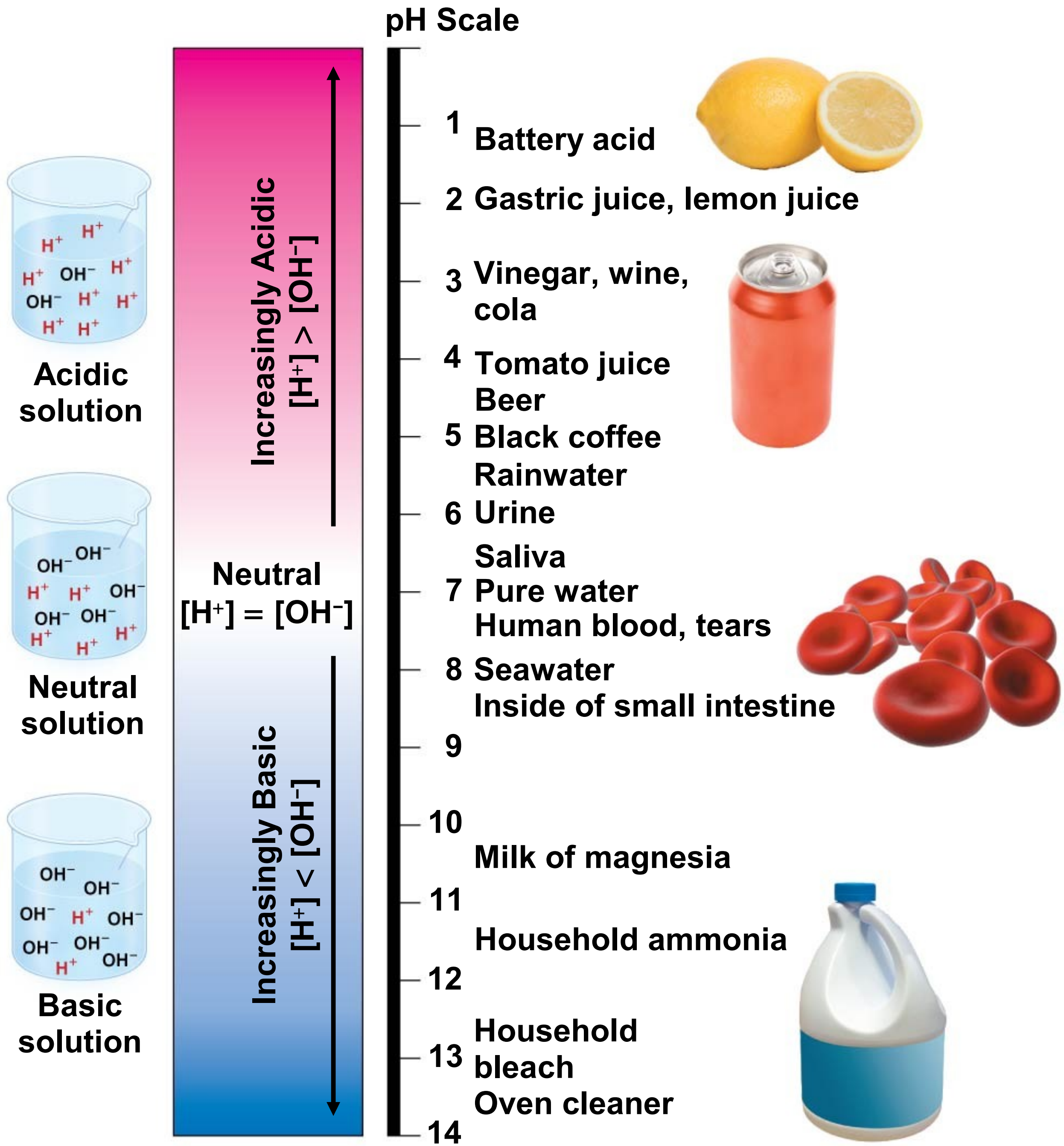
$$pH = -\log [H^+]$$

- For a neutral aqueous solution, $[H^+]$ is 10^{-7}, so

$$-\log [H^+] = -(-7) = 7$$

- Acidic solutions have pH values less than 7
- Basic solutions have pH values greater than 7
- Most biological fluids have pH values in the range of 6 to 8

Figure 2.23



pH Scale
Acidic solution
Neutral solution
Basic solution
Increasingly Acidic
[H+] > [OH–]
Neutral
[H+] = [OH–]
Increasingly Basic
[H+] < [OH–]
1 Battery acid
2 Gastric juice, lemon juice
3 Vinegar, wine, cola
4 Tomato juice
Beer
5 Black coffee
Rainwater
6 Urine
Saliva
7 Pure water
Human blood, tears
8 Seawater
Inside of small intestine
9
10
Milk of magnesia
11
Household ammonia
12
Household bleach
13
Oven cleaner
14

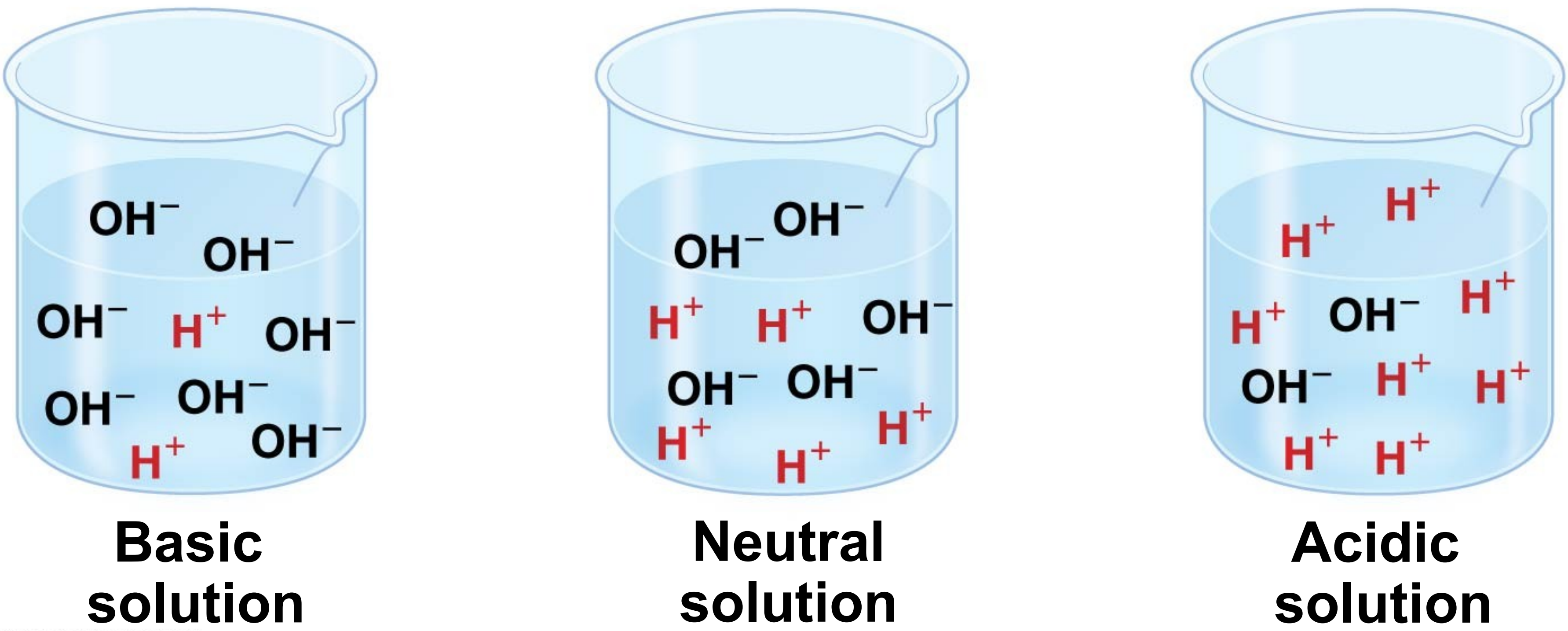
Figure 2.23a

Figure 2.23b

Figure 2.23c

Figure 2.23d

Figure 2.23e


OH$^-$
OH$^-$
OH$^-$ H$^+$ OH$^-$
OH$^-$ OH$^-$
H$^+$ OH$^-$
Basic solution
OH$^-$ OH$^-$
H$^+$ H$^+$ OH$^-$
OH$^-$ OH$^-$
H$^+$ H$^+$ H$^+$
Neutral solution
H$^+$
H$^+$
H$^+$ OH$^-$ H$^+$
OH$^-$ H$^+$ H$^+$
H$^+$ H$^+$
Acidic solution

Buffers

- The internal pH of most living cells must remain close to pH 7
- **Buffers** are substances that minimize changes in concentrations of H^+ and OH^- in a solution
- Most buffers consist of an acid-base pair that reversibly combines with H^+

- Carbonic acid is a buffer that contributes to pH stability in human blood:

Response to a rise in pH

$$H_2CO_3 \rightleftharpoons HCO_3^- + H^+$$

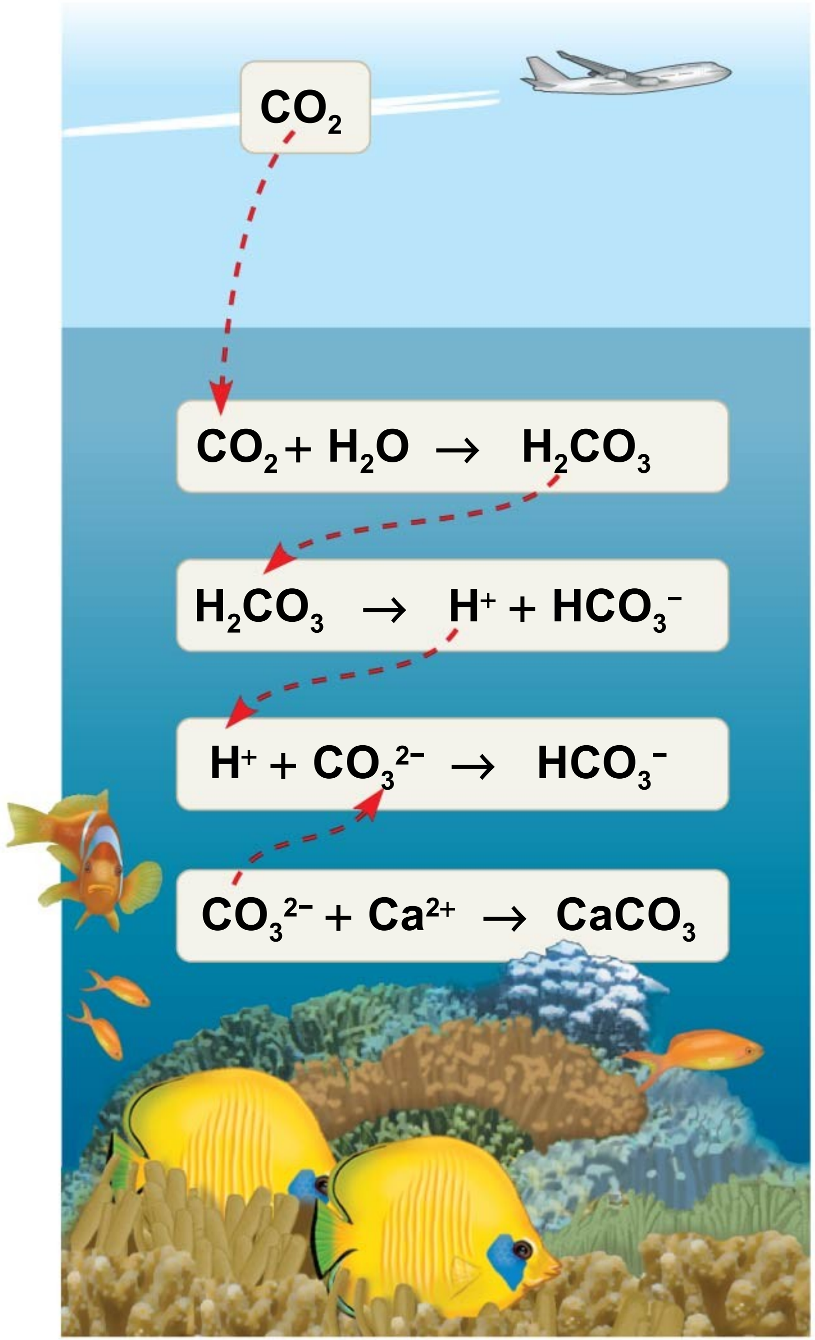
H^+ donor (acid) — Response to a drop in pH — H^+ acceptor (base) — Hydrogen ion

Acidification: A Threat to Our Oceans

- Human activities such as burning fossil fuels threaten water quality
- CO_2 is the main product of fossil fuel combustion
- About 25% of human-generated CO_2 is absorbed by the oceans
- CO_2 dissolved in seawater forms carbonic acid; this causes ocean acidification

- As seawater acidifies, H^+ ions combine with CO_3^{2-} ions to form bicarbonate ions (HCO_3^-)
- It is predicted that carbonate ion concentrations will decline by 40% by the year 2100
- This is a concern because organisms that build coral reefs or shells require carbonate ions

Figure 2.24



CO_2
$CO_2 + H_2O \rightarrow H_2CO_3$
$H_2CO_3 \rightarrow H^+ + HCO_3^-$
$H^+ + CO_3^{2-} \rightarrow HCO_3^-$
$CO_3^{2-} + Ca^{2+} \rightarrow CaCO_3$

Figure 2.UN04

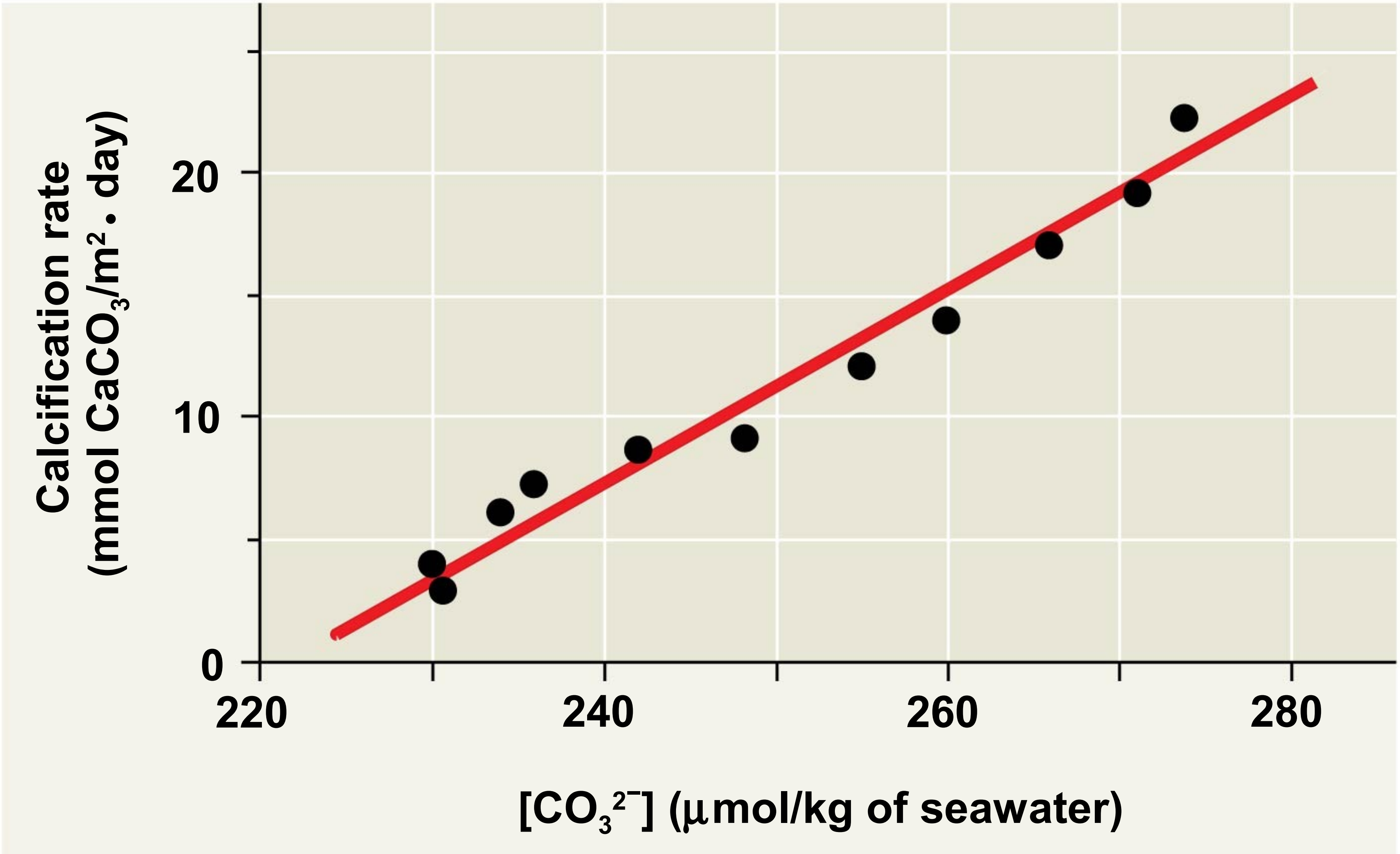


Figure 2.UN05

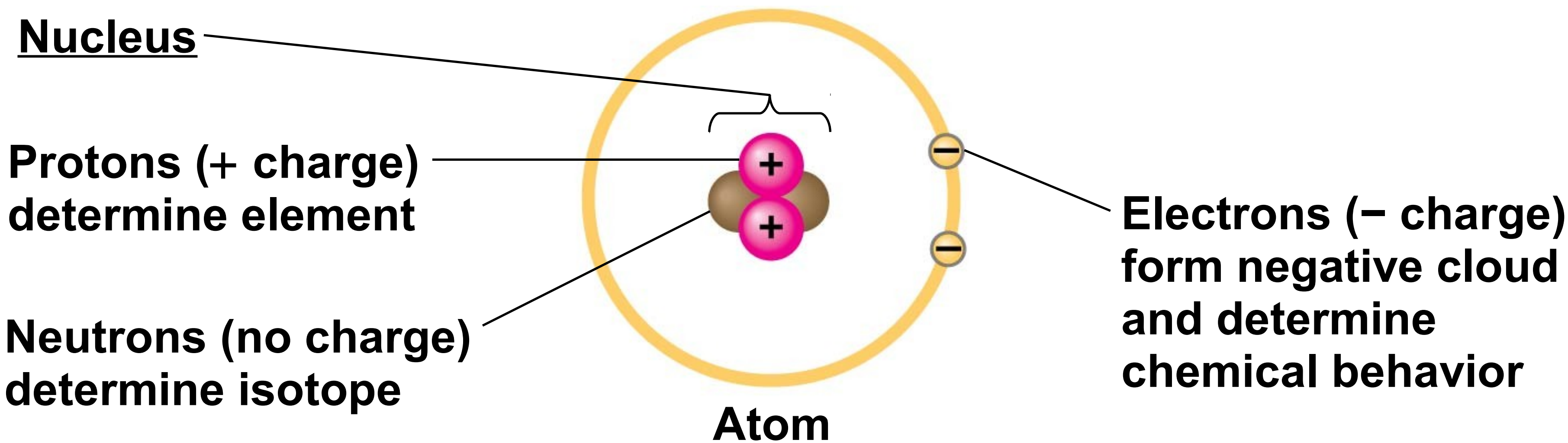


Figure 2.UN06

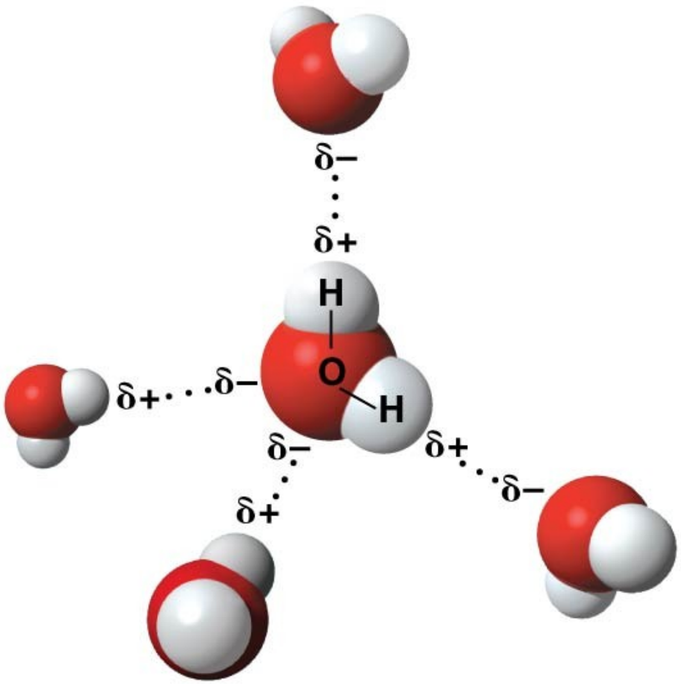
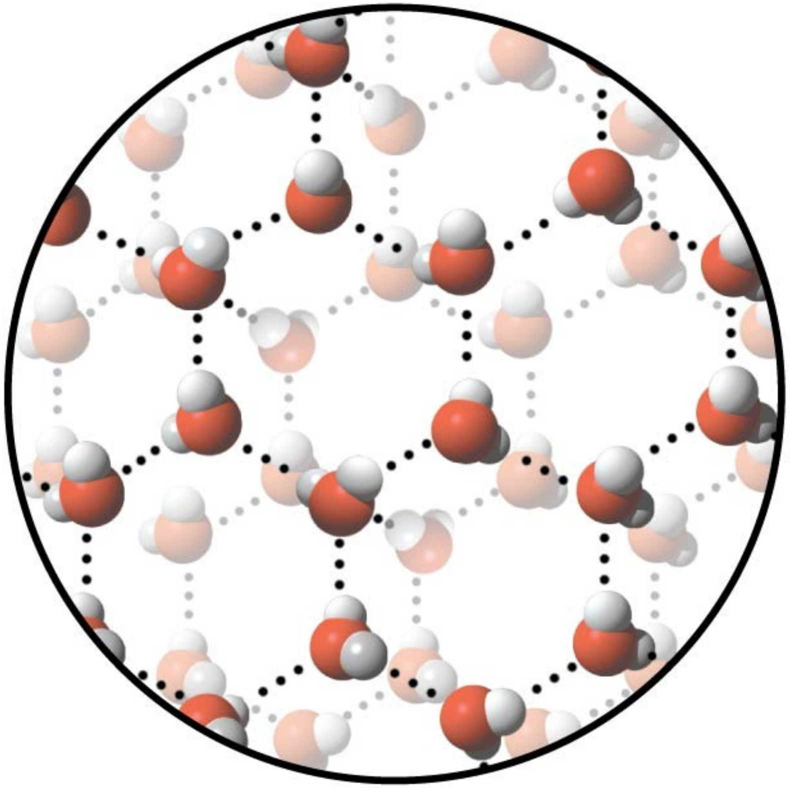
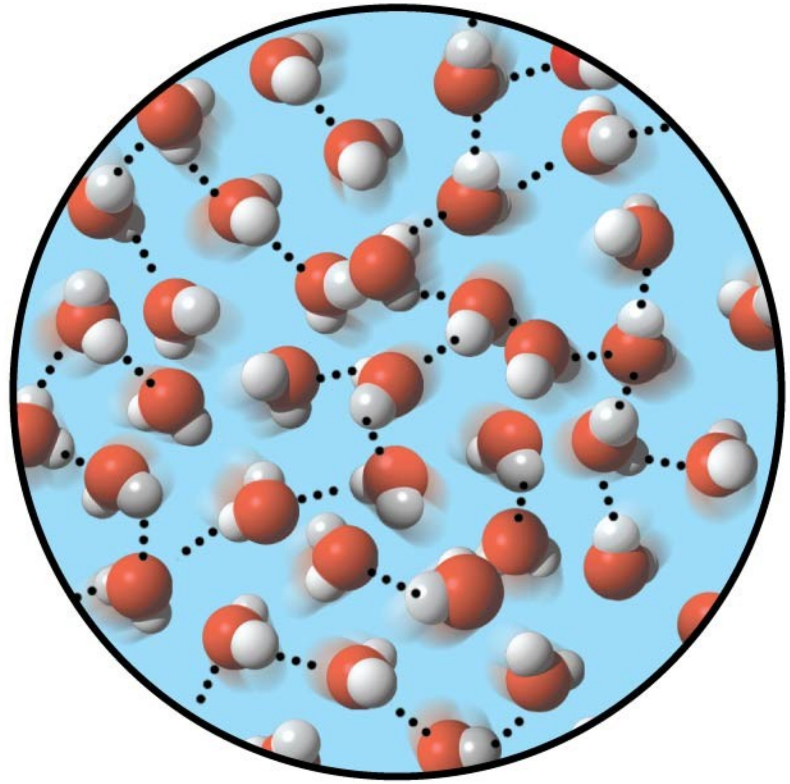


Figure 2.UN07



Ice: stable hydrogen bonds



Liquid water: transient hydrogen bonds

Figure 2.UN08

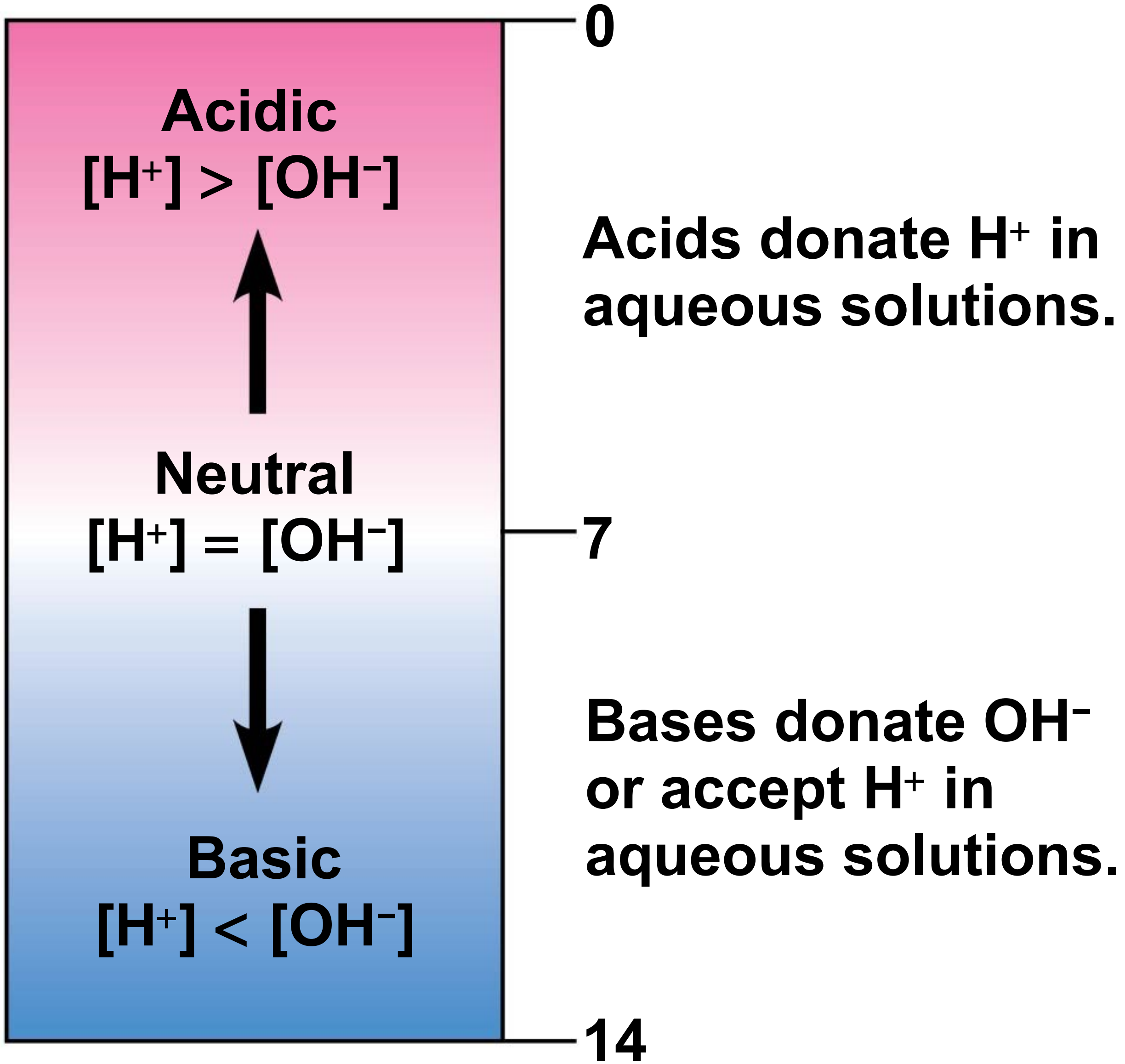


Figure 2.UN09