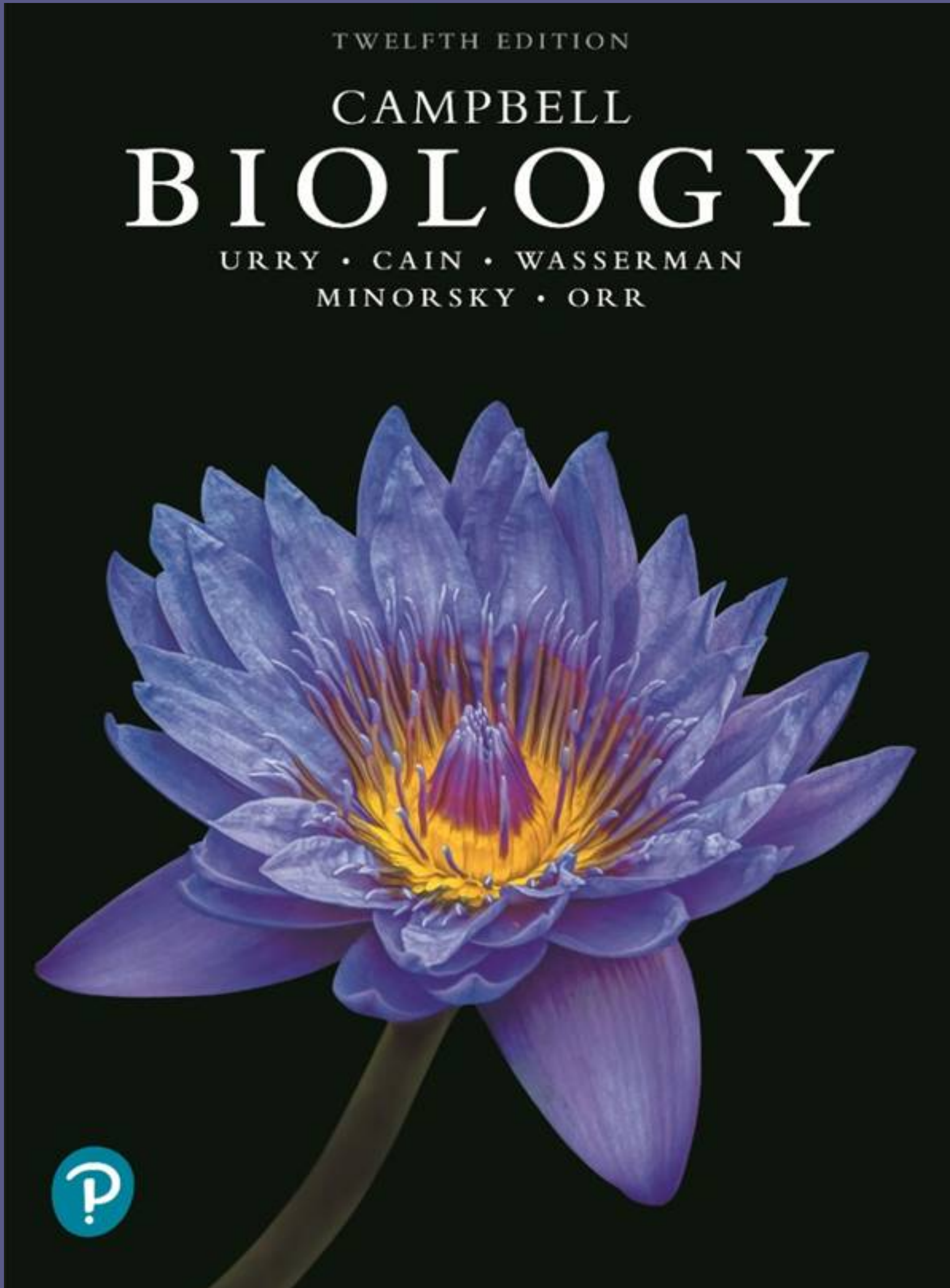



# Chapter 2

# The Chemical Context of Life

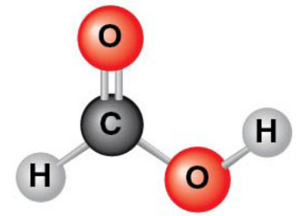
Lecture Presentations by
Nicole Tunbridge and
Kathleen Fitzpatrick

Figure 2.1

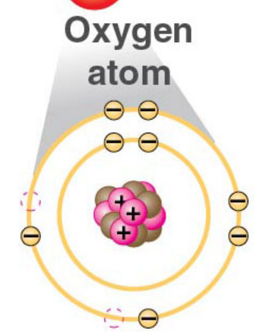
## What determines the properties of a compound such as formic acid?

A **compound** is made of atoms joined by bonds.

Formic acid

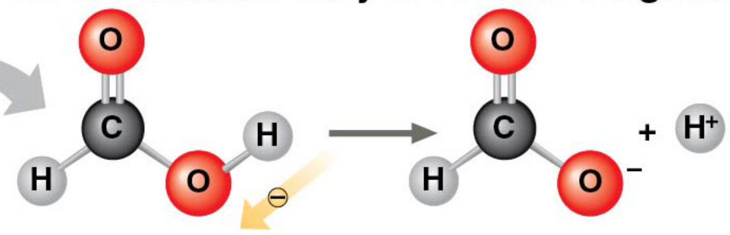



The **number of protons (⊕)** determines an atom's identity.




An atom's **electron (⊖) distribution** determines its ability to form bonds.

A **compound's properties** depend on its atoms and how they are bonded together.



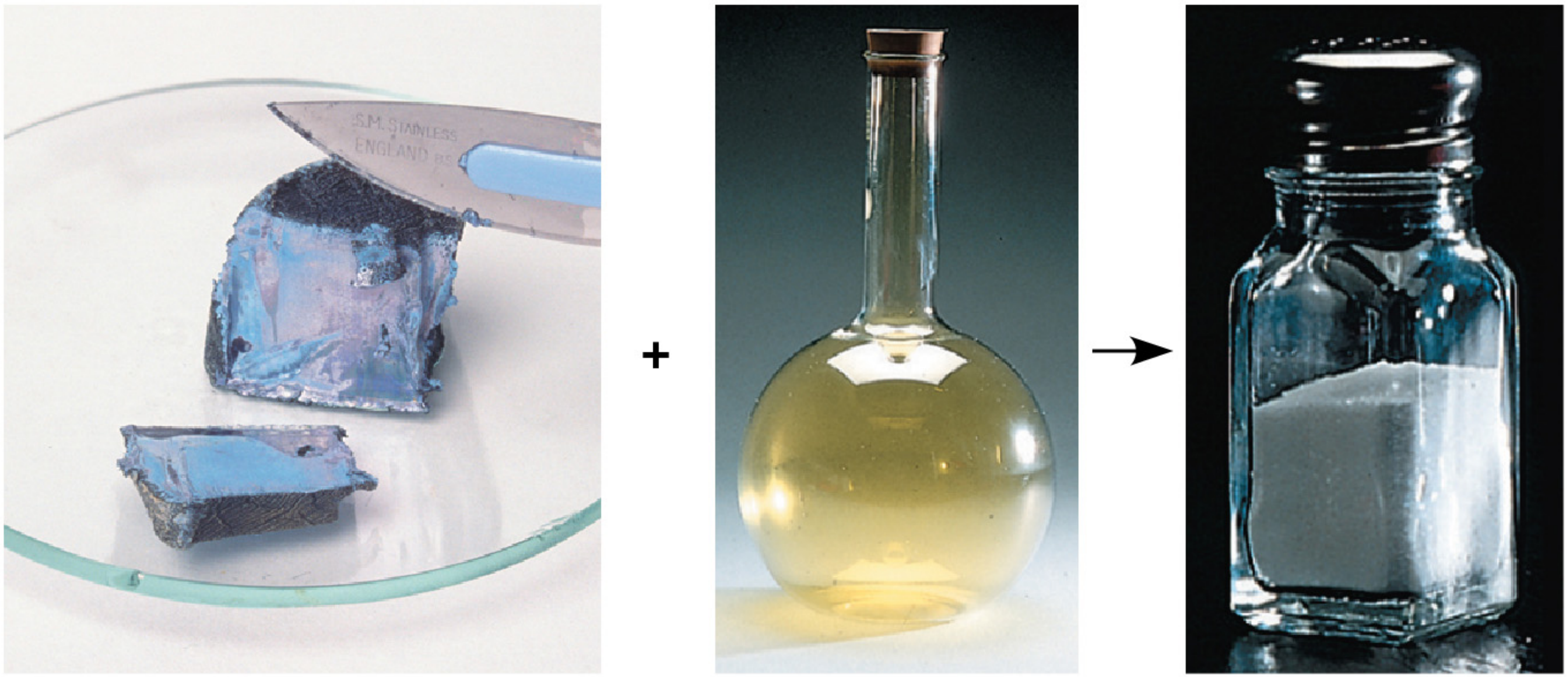
## CONCEPT 2.1: Matter consists of chemical elements in pure form and in combinations called compounds

- 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 characteristics (emergent properties) different from those of its elements

Figure 2.2




Na + Cl → NaCl

Sodium + Chlorine (gas) → Sodium chloride

# The Elements of Life

- About 20–25% of the 92 natural elements are required for life (**essential elements**)
- Carbon, hydrogen, oxygen, and nitrogen make up 96% of living matter
- Most of the remaining 4% consists of calcium, phosphorus, potassium, and sulfur
- **Trace elements** are required by an organism in only minute quantities

**Table 2.1 Elements in the Human Body**

| Element | Symbol | Percentage of Body Mass (including water) | |
|---|---|---|---|
| Oxygen | O | 65.0% | 96.3% |
| Carbon | C | 18.5% | |
| Hydrogen | H | 9.5% | |
| Nitrogen | N | 3.3% | |
| Calcium | Ca | 1.5% | 3.7% |
| Phosphorus | P | 1.0% | |
| Potassium | K | 0.4% | |
| Sulfur | S | 0.3% | |
| Sodium | Na | 0.2% | |
| Chlorine | Cl | 0.2% | |
| Magnesium | Mg | 0.1% | |
| Trace elements (less than 0.01% of mass): Boron (B), chromium (Cr), cobalt (Co), copper (Cu), fluorine (F), iodine (I), iron (Fe), manganese (Mn), molybdenum (Mo), selenium (Se), silicon (Si), tin (Sn), vanadium (V), zinc (Zn) | | | |

# *Case Study*: Evolution of Tolerance to Toxic Elements

- Some elements can be toxic
- Some species can become adapted to environments containing toxic elements
  - For example, some plant communities are adapted to serpentine

Figure 2.3

## CONCEPT 2.2: An element’s properties depend on the structure of its atoms

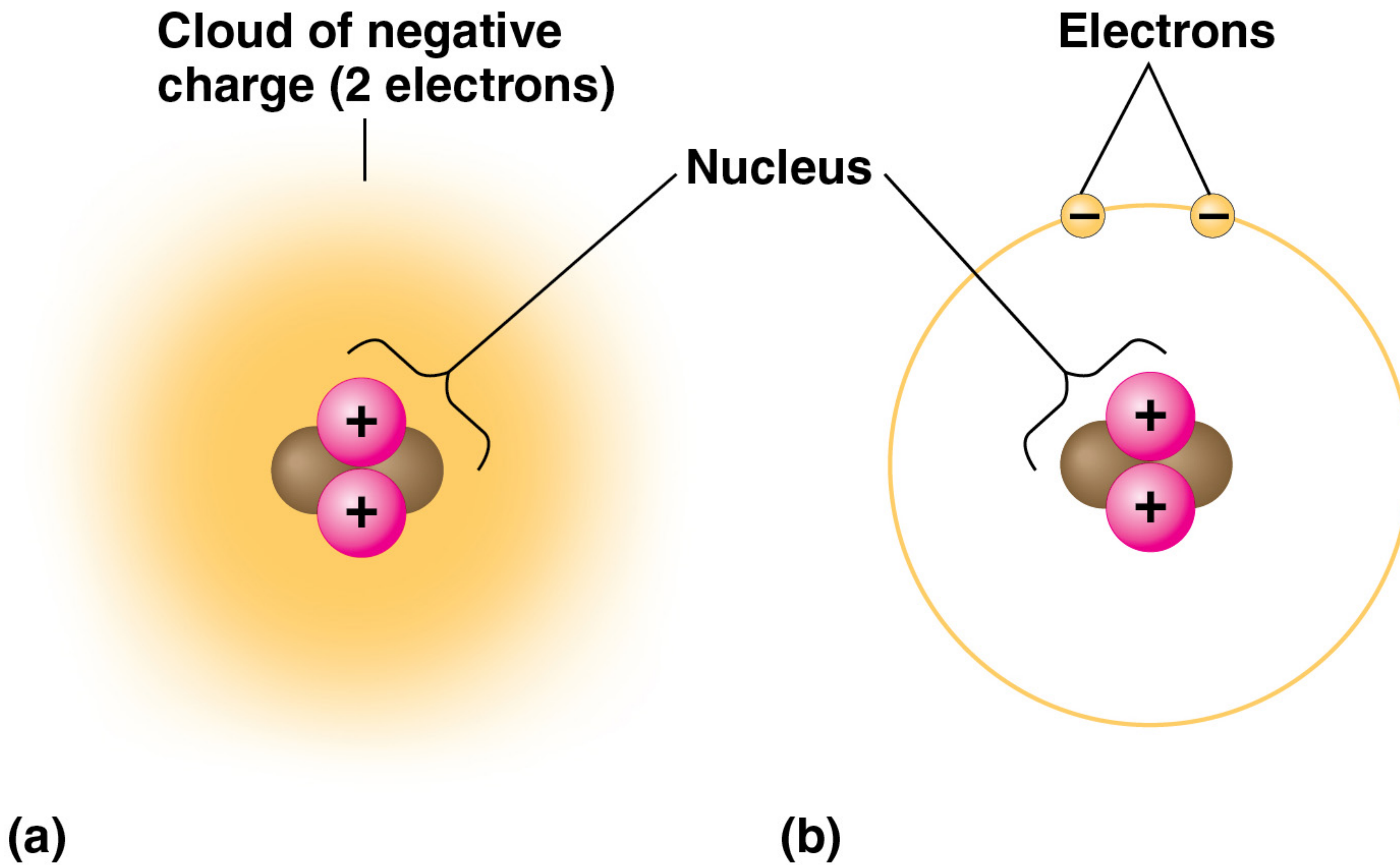
- Each element consists of unique atoms
- An **atom** is the smallest unit of matter that still retains the properties of an element

# Subatomic Particles

- Atoms are composed of 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” of negative charge around the nucleus
- Neutron mass and proton mass are almost identical and are measured in **daltons**
- Electrons are so small they are ignored when calculating the total mass of an atom

Figure 2.4



# Atomic Number and Atomic Mass

- Atoms of the various elements differ in their 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

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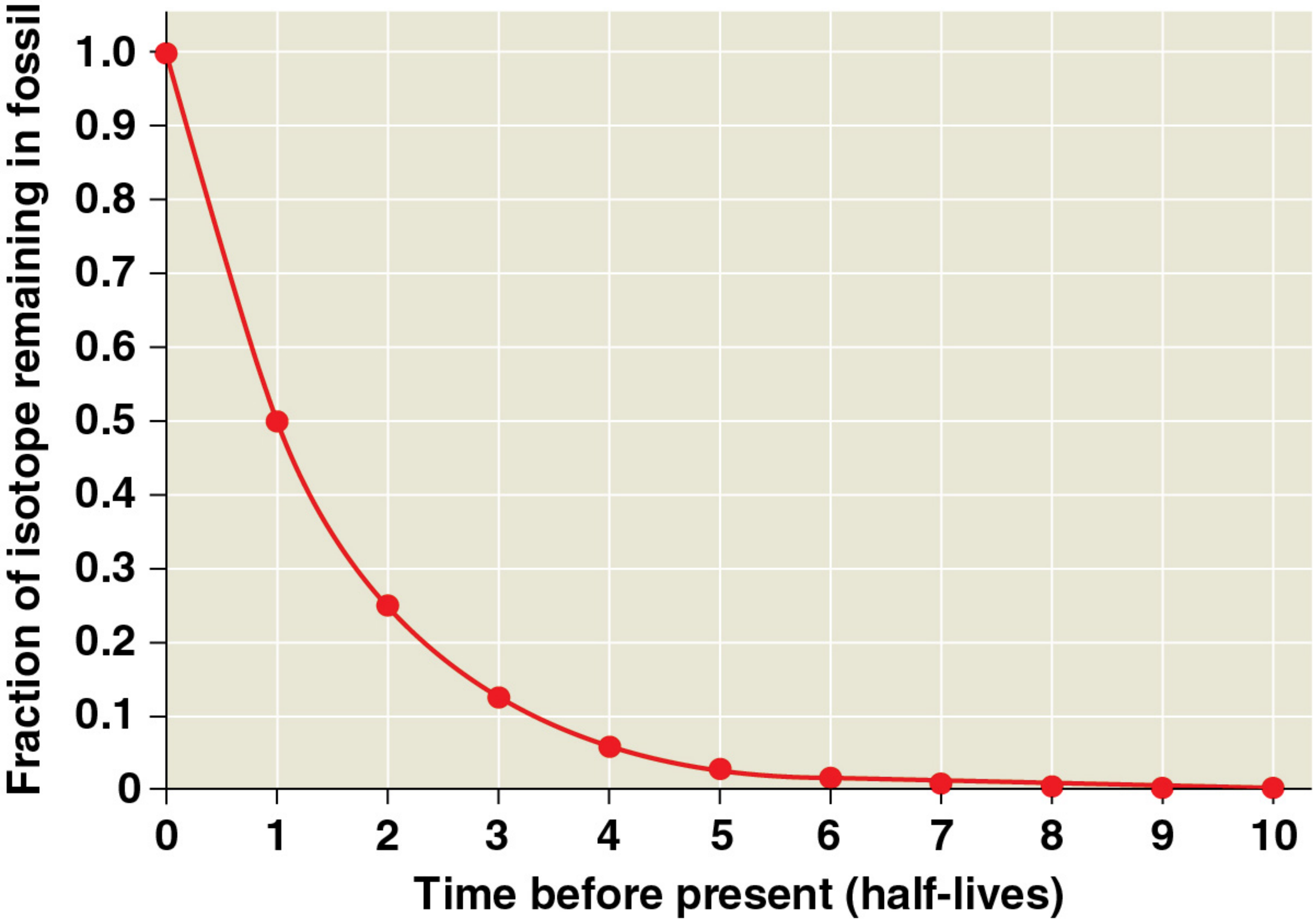
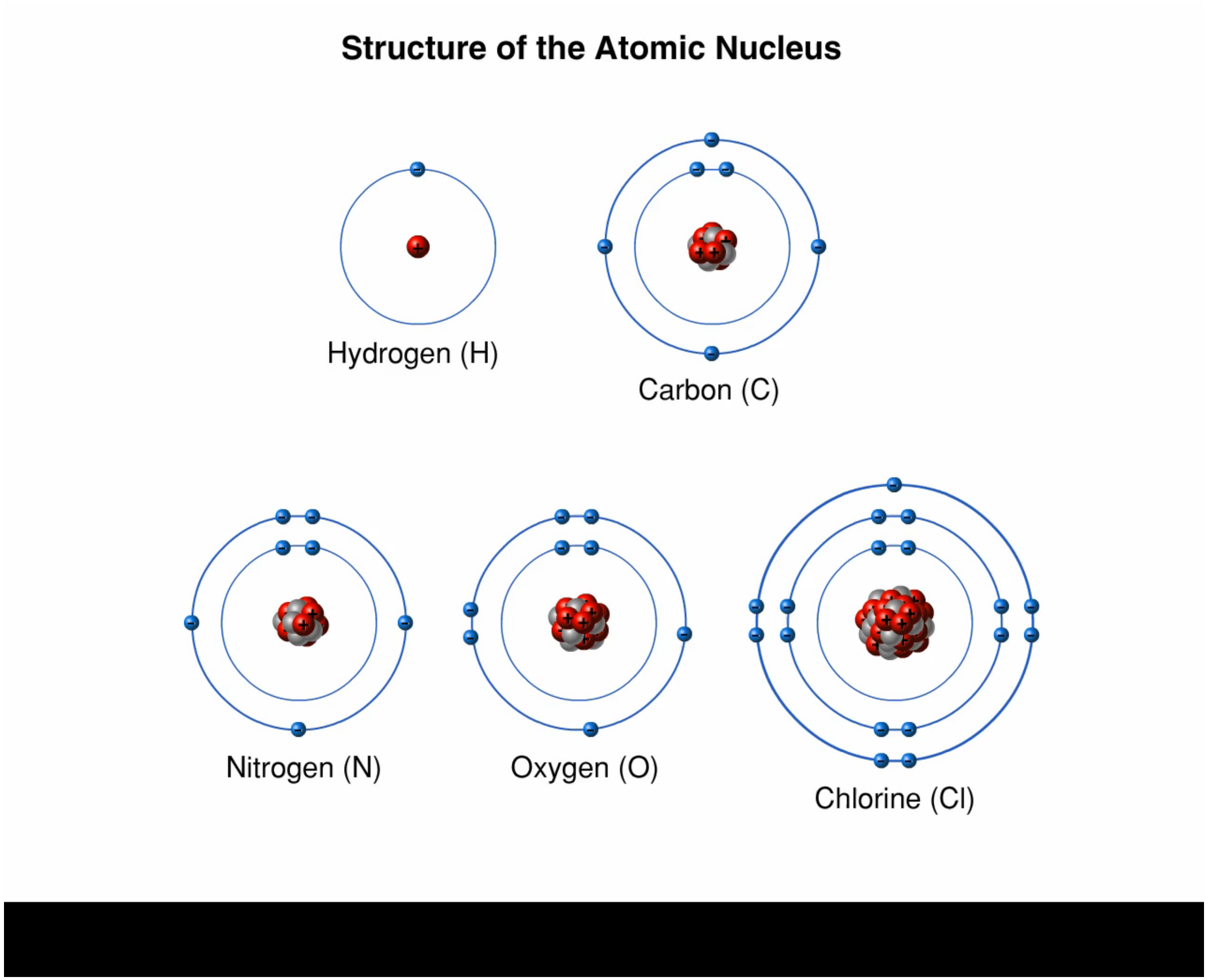



**Data from R. Pinhasi et al., Revised age of late Neanderthal occupation and the end of the Middle Paleolithic in the northern Caucasus, *Proceedings of the National Academy of Sciences USA* 147:8611–8616 (2011). doi 10.1073/pnas.1018938108**

# Animation: Atomic Number and Atomic Mass



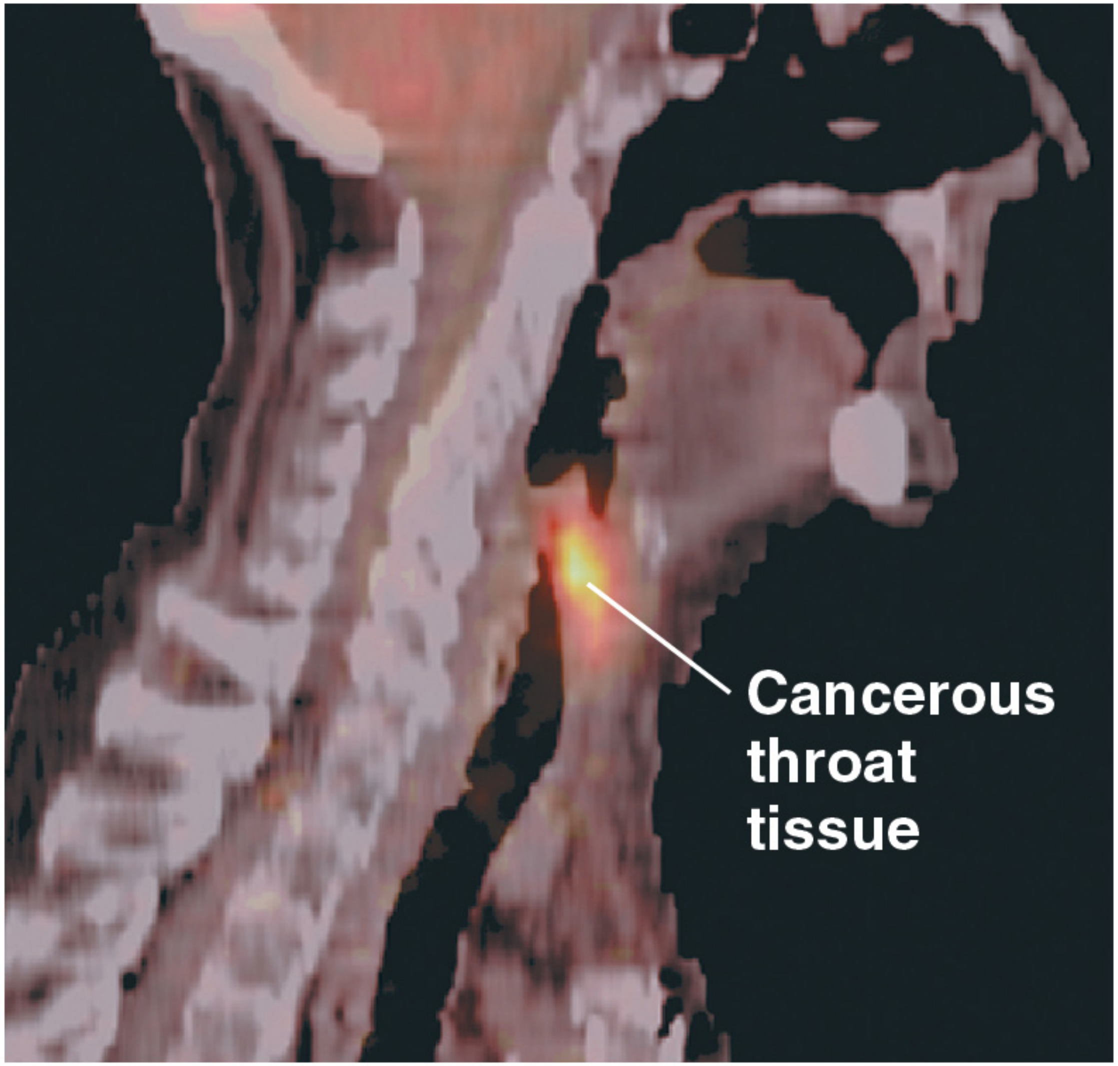
# Isotopes

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

# *Radioactive Tracers*

- Radioactive isotopes are often used as diagnostic tools in medicine
- Radioactive tracers can be used to track atoms through metabolism
- They can also be used in combination with sophisticated imaging instruments
- PET scanners can monitor the growth and metabolism of cancers in the body

Figure 2.5



# *Radiometric Dating*

- A “parent” isotope decays into its “daughter” isotope at a fixed rate, expressed as the **half-life** of the isotope
- In **radiometric dating**, scientists measure the ratio of different isotopes and calculate how many half-lives have passed since the fossil or rock was formed
- Half-life values vary from seconds or days for some isotopes to billions of years for others

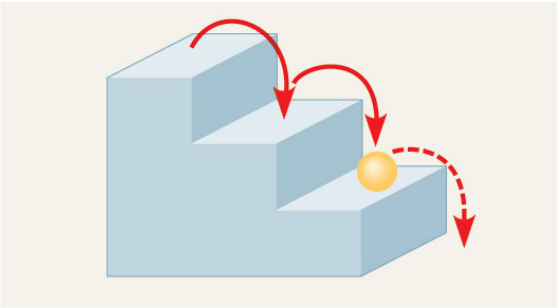
# The Energy Levels of Electrons

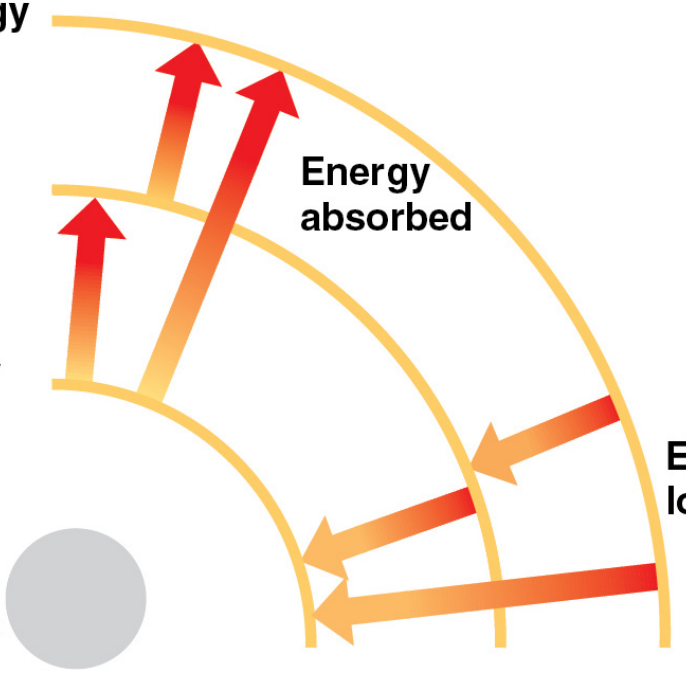
- **Energy** is the capacity to cause change
- **Potential energy** is the energy that matter possesses because of its location or structure
- Matter has a natural tendency to move toward the lowest possible state of potential energy

- The electrons of an atom differ in their amounts of potential energy based on their distance from the nucleus
- Changes in potential energy of electrons can occur only in steps of fixed amounts
- Electrons are found in different **electron shells**, each with a characteristic average distance and energy level

Figure 2.6

(a) A ball bouncing down a flight of stairs can come to rest only on each step, not between steps.



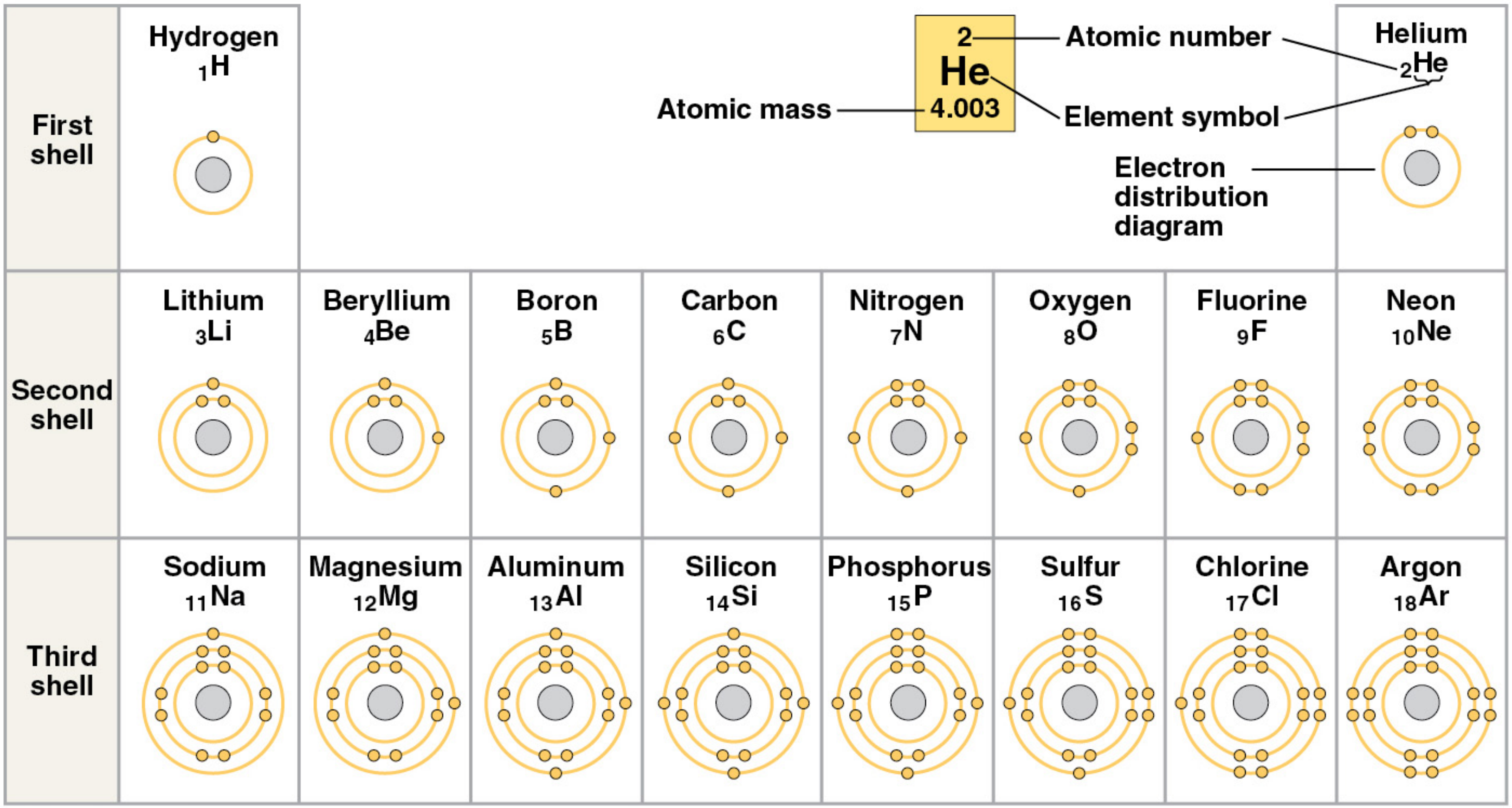



(b)

# Electron Distribution and Chemical Properties

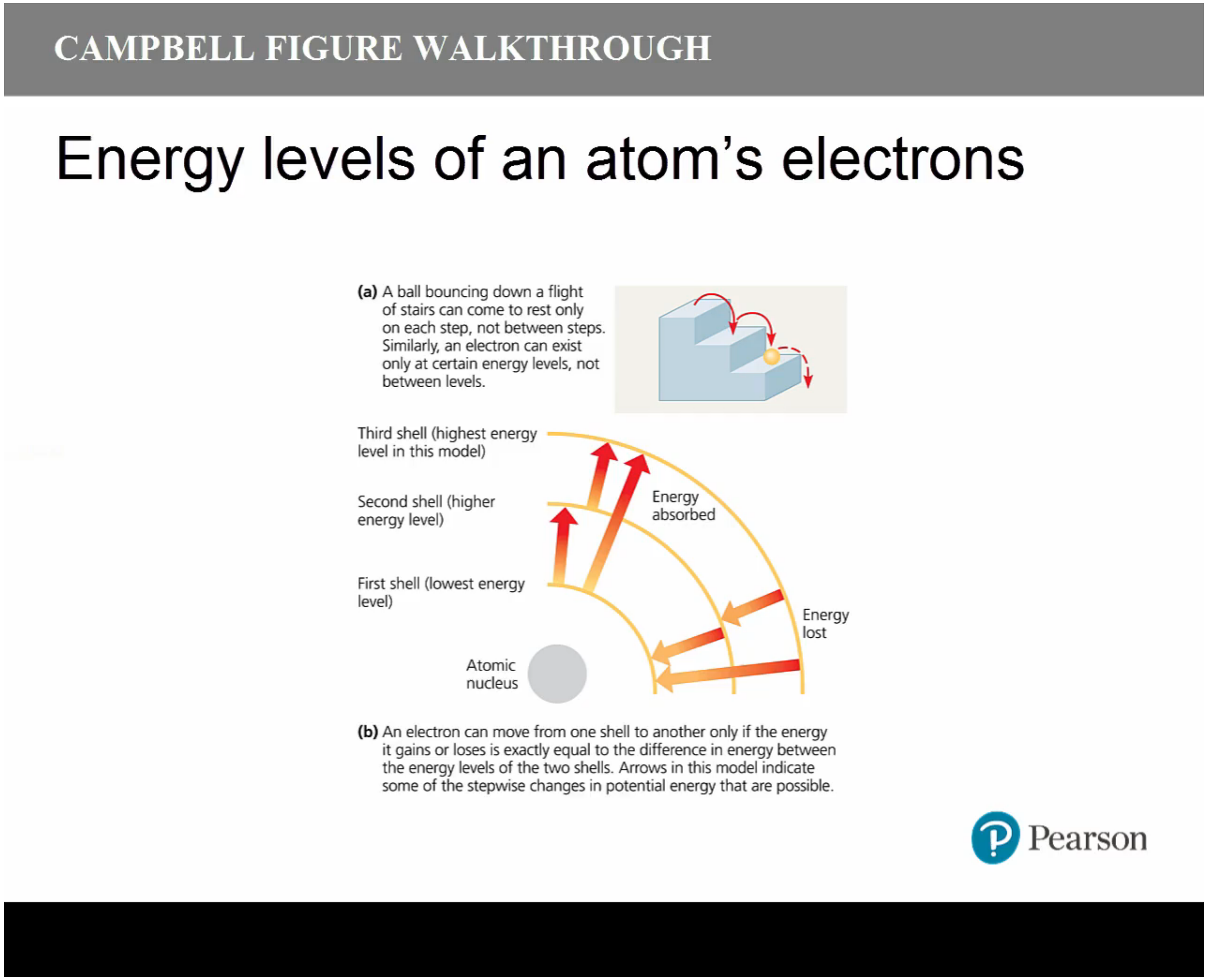
- The chemical behavior of an atom is determined by the distribution of electrons in the electron shells
- The periodic table of the elements shows the electron distribution for each element
- The left-to-right sequence of elements in each row corresponds to the sequential addition of electrons and protons

Figure 2.7

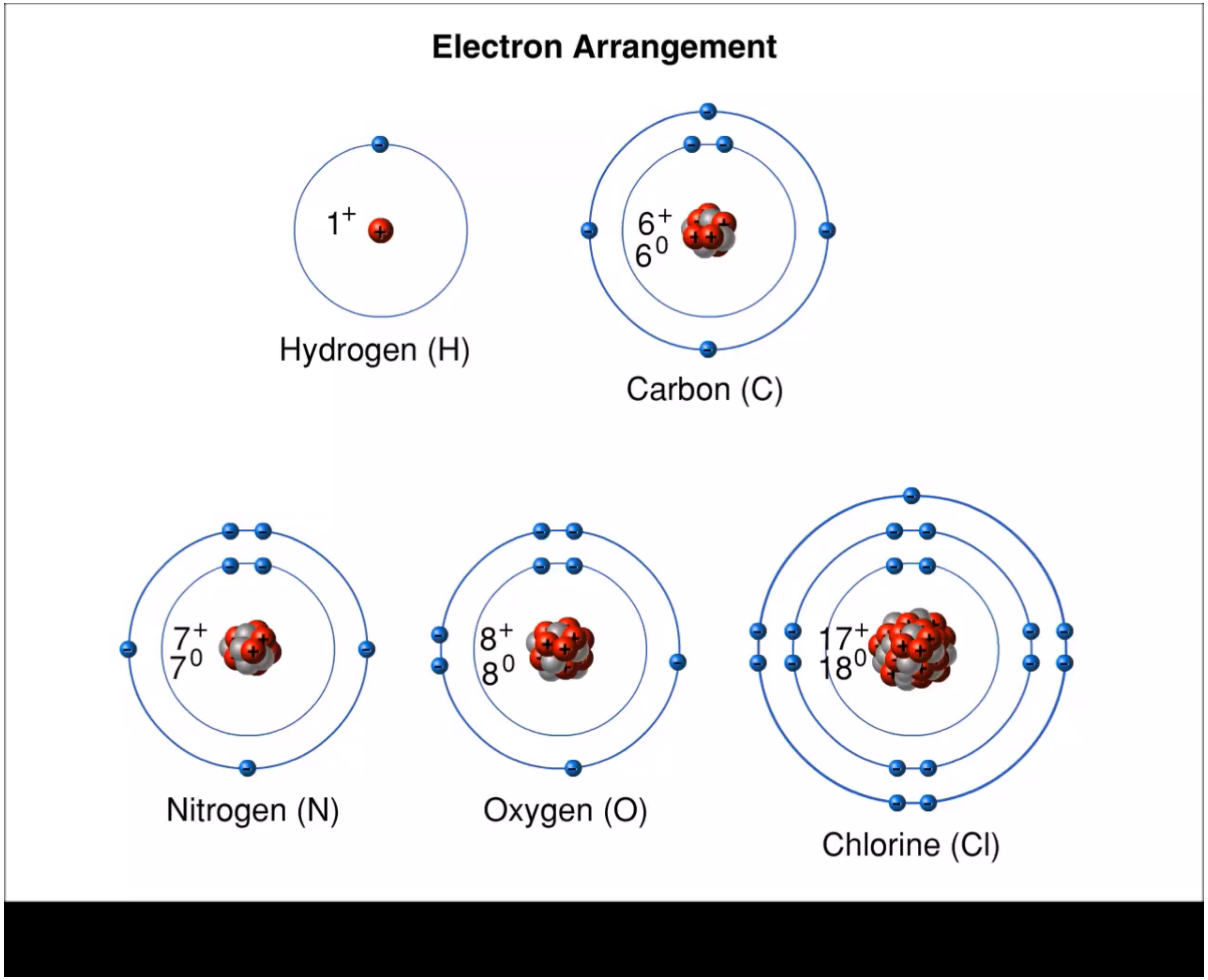


First shell
Second shell
Third shell
Hydrogen
1H
2
He
4.003
Atomic number
Atomic mass
Element symbol
Electron distribution diagram
Helium
2He
Lithium
3Li
Beryllium
4Be
Boron
5B
Carbon
6C
Nitrogen
7N
Oxygen
8O
Fluorine
9F
Neon
10Ne
Sodium
11Na
Magnesium
12Mg
Aluminum
13Al
Silicon
14Si
Phosphorus
15P
Sulfur
16S
Chlorine
17Cl
Argon
18Ar

# Animation: Energy Levels Atoms Electrons



# Animation: Electron Distribution Diagrams

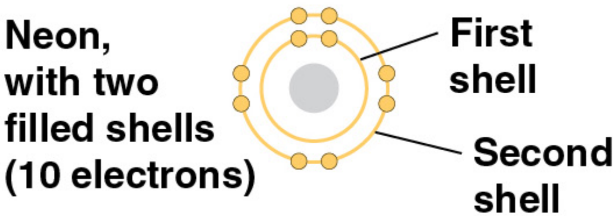


- **Valence electrons** are those in the outermost shell, or **valence shell**
- The chemical behavior of an atom is mostly determined by the number of valence electrons
- Elements with a full valence shell are chemically inert

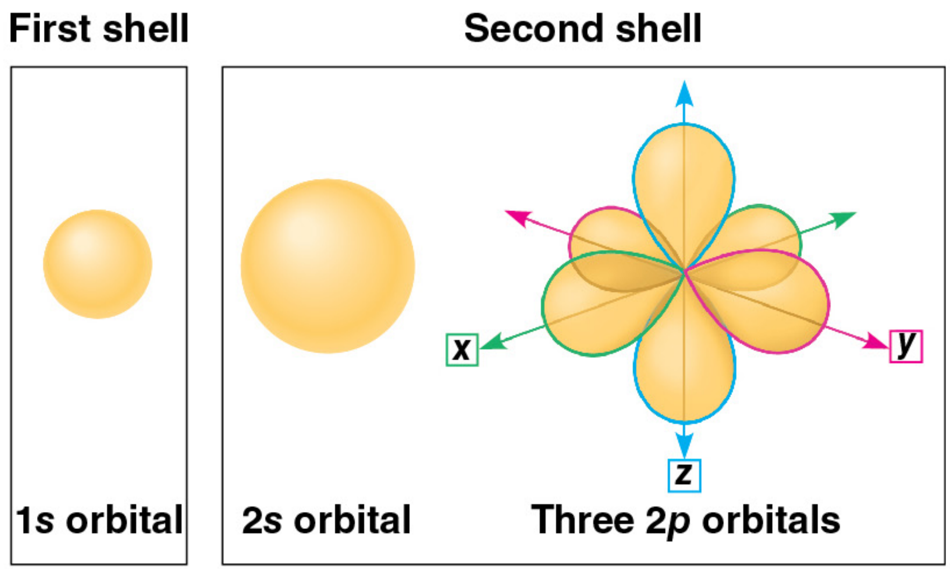
# Electron Orbitals

- An **orbital** is the three-dimensional space where an electron is found 90% of the time
- Each electron shell consists of a specific number of orbitals
- No more than 2 electrons can occupy a single orbital
- Atoms interact in a way that completes their valence shells

Figure 2.8




**(a) Electron distribution diagram**




**(b) Separate electron orbitals**


**(c) Superimposed electron orbitals**

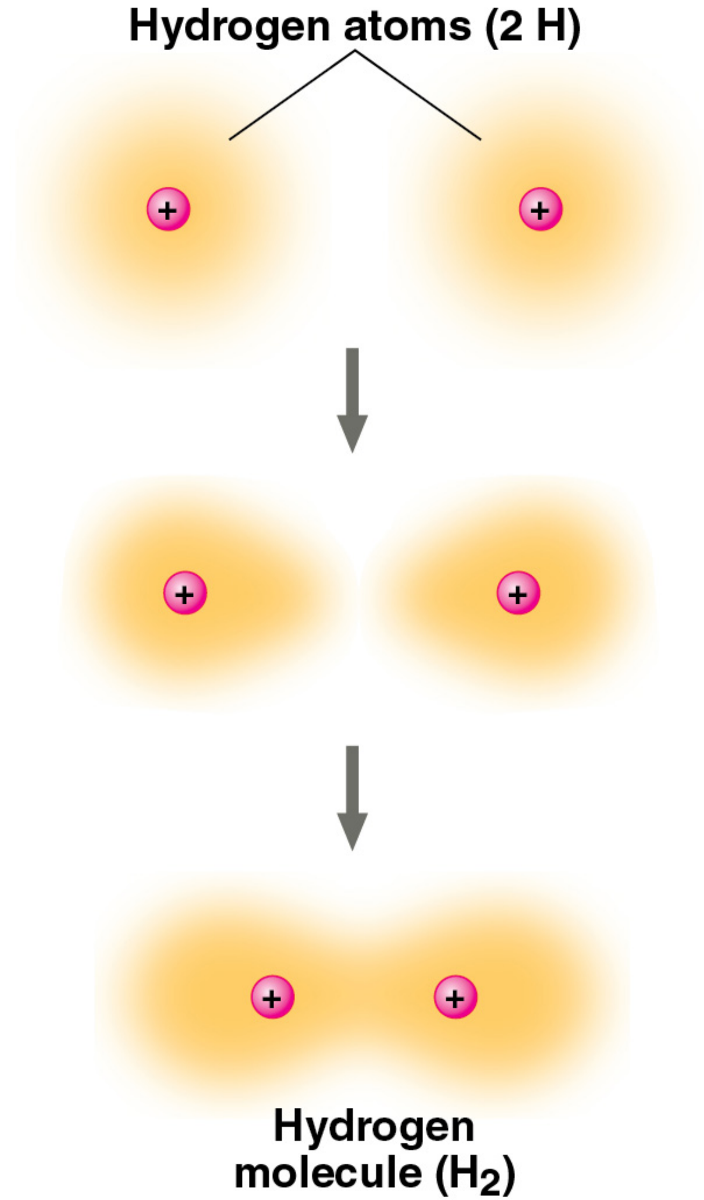
# CONCEPT 2.3: The formation and function of molecules and ionic compounds depend on chemical bonding between atoms

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

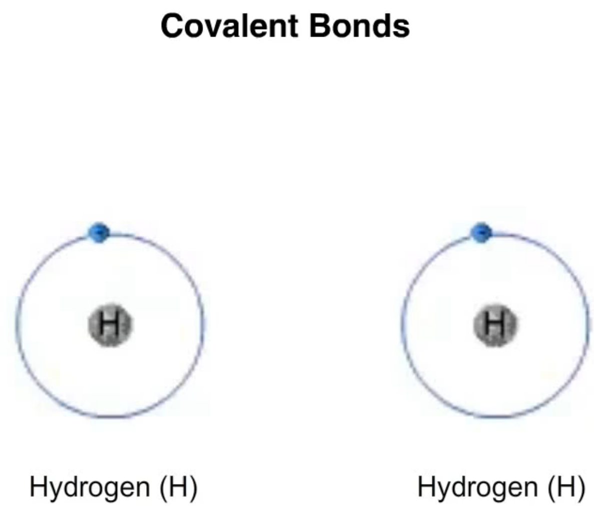
# Covalent Bonds

- 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

Figure 2.9



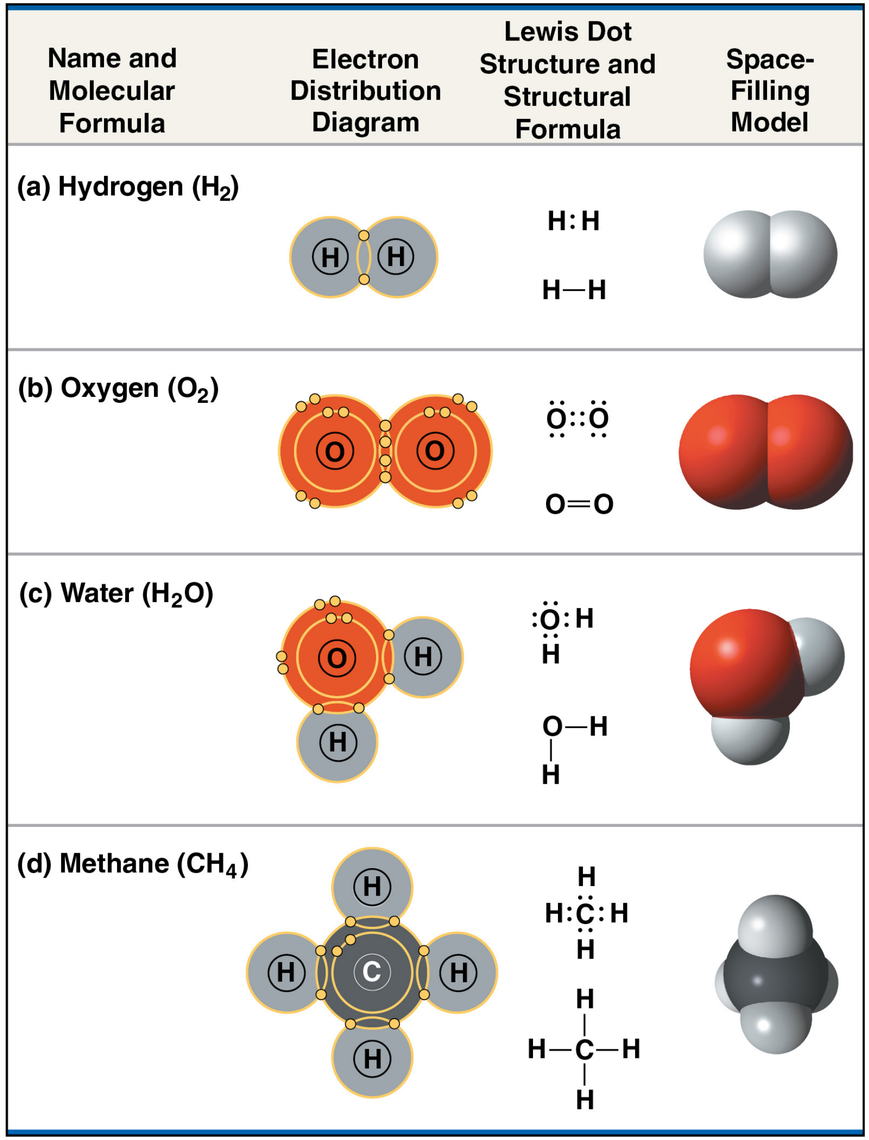
# Animation: Covalent Bonds



- A **molecule** consists of two or more atoms held together by covalent bonds
- A single covalent bond, or **single bond**, is the sharing of one pair of valence electrons
- A double covalent bond, or **double bond**, is the sharing of two pairs of valence electrons

- The notation used to represent atoms and bonding is called a structural formula
  - For example, H—H represents a single bond
  - O ═ O represents a double bond
- This can be abbreviated further with a molecular formula
  - For example, $H_2$

Figure 2.10

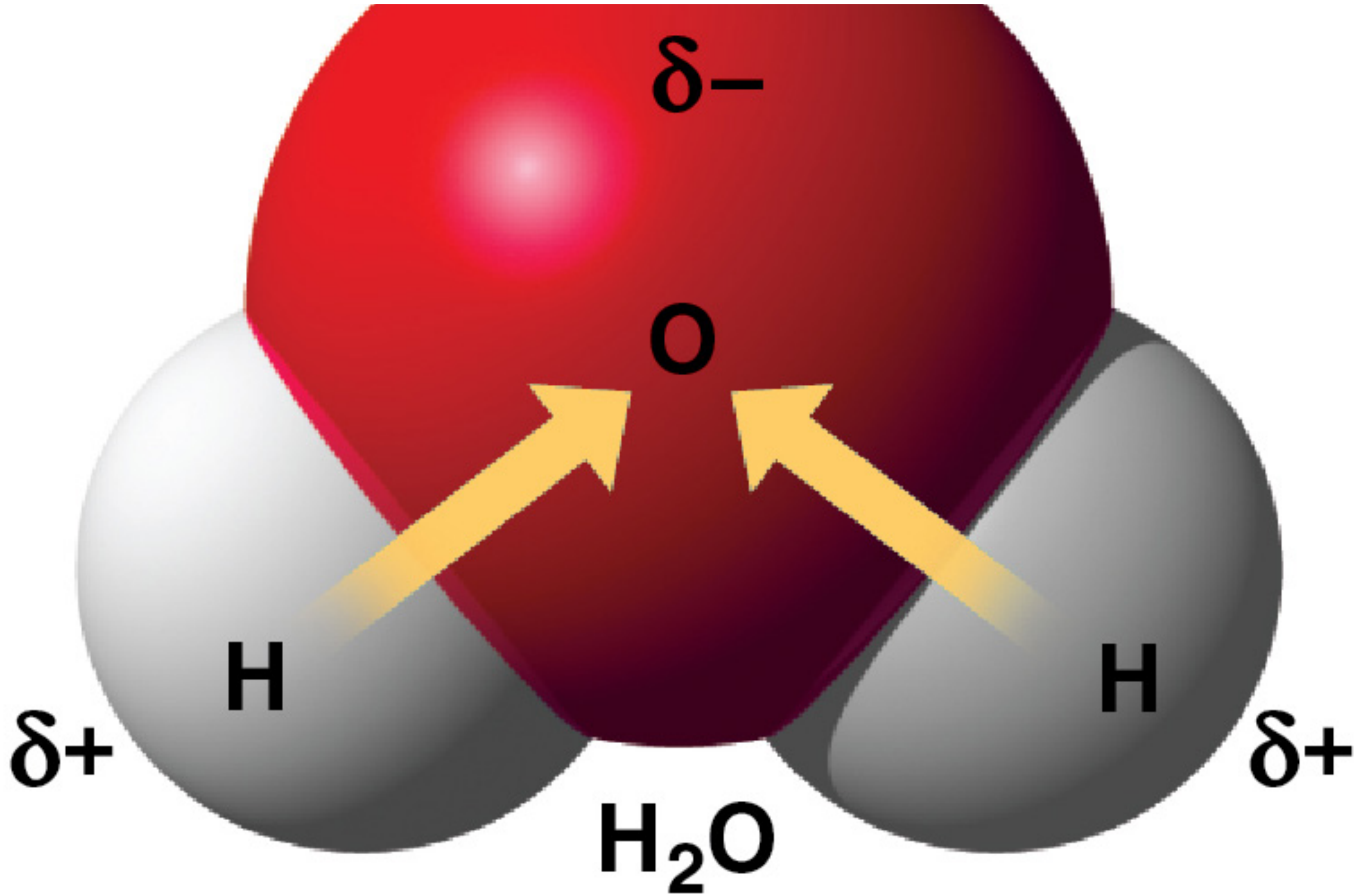

Name and Molecular Formula
Electron Distribution Diagram
Lewis Dot Structure and Structural Formula
Space-Filling Model
(a) Hydrogen ($H_2$)
H
H:H
H—H
(b) Oxygen ($O_2$)
O
O=O
(c) Water ($H_2O$)
O
H
O—H
H
(d) Methane ($CH_4$)
C
H
H—C—H
H

- Bonding capacity is called the atom’s **valence**
- Covalent bonds can form between atoms of the same element or atoms of different elements
- A compound is a combination of two or more different elements

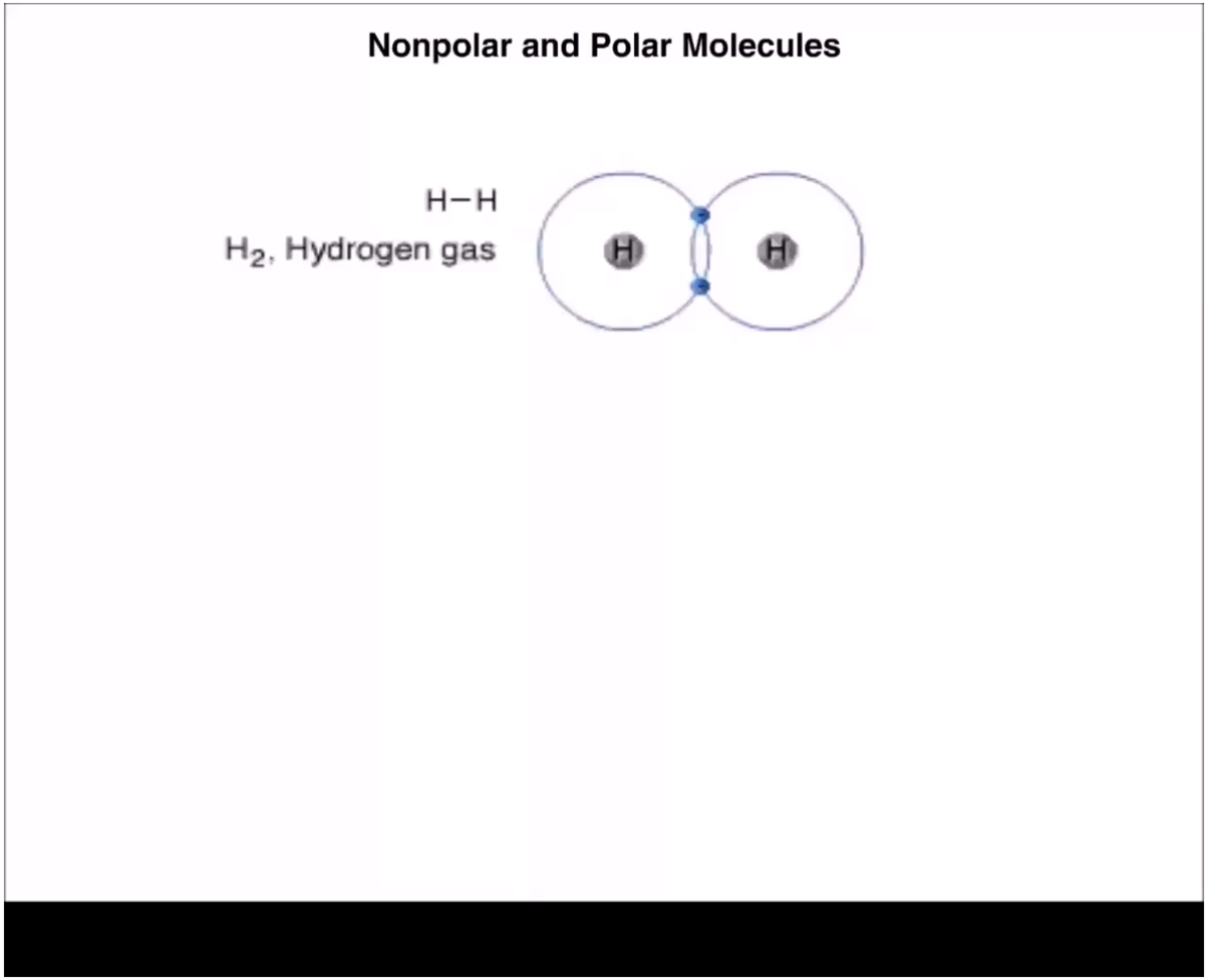
- Atoms in a molecule attract electrons to varying degrees
- **Electronegativity** is an atom’s attraction for the electrons in a covalent bond
- The more electronegative an atom is, the more strongly it pulls shared electrons toward itself

- 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

Figure 2.11



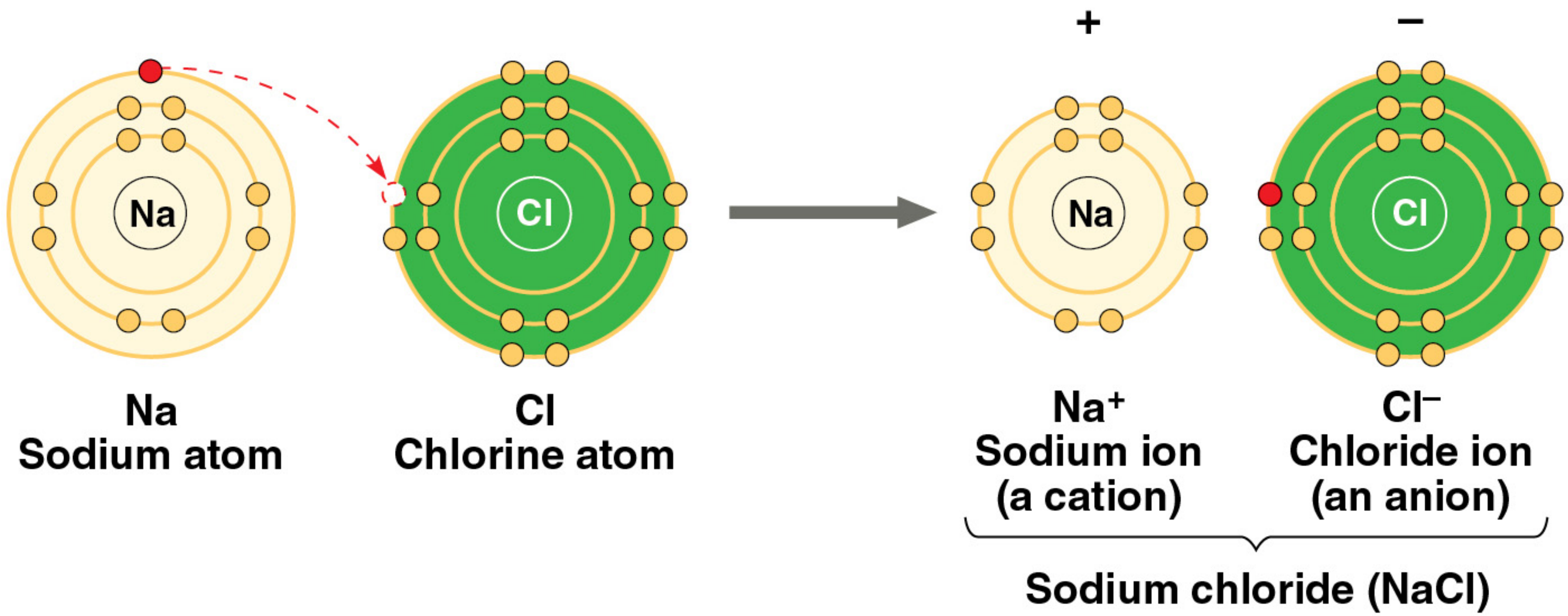
# Animation: Nonpolar And Polar Molecules



# Ionic Bonds

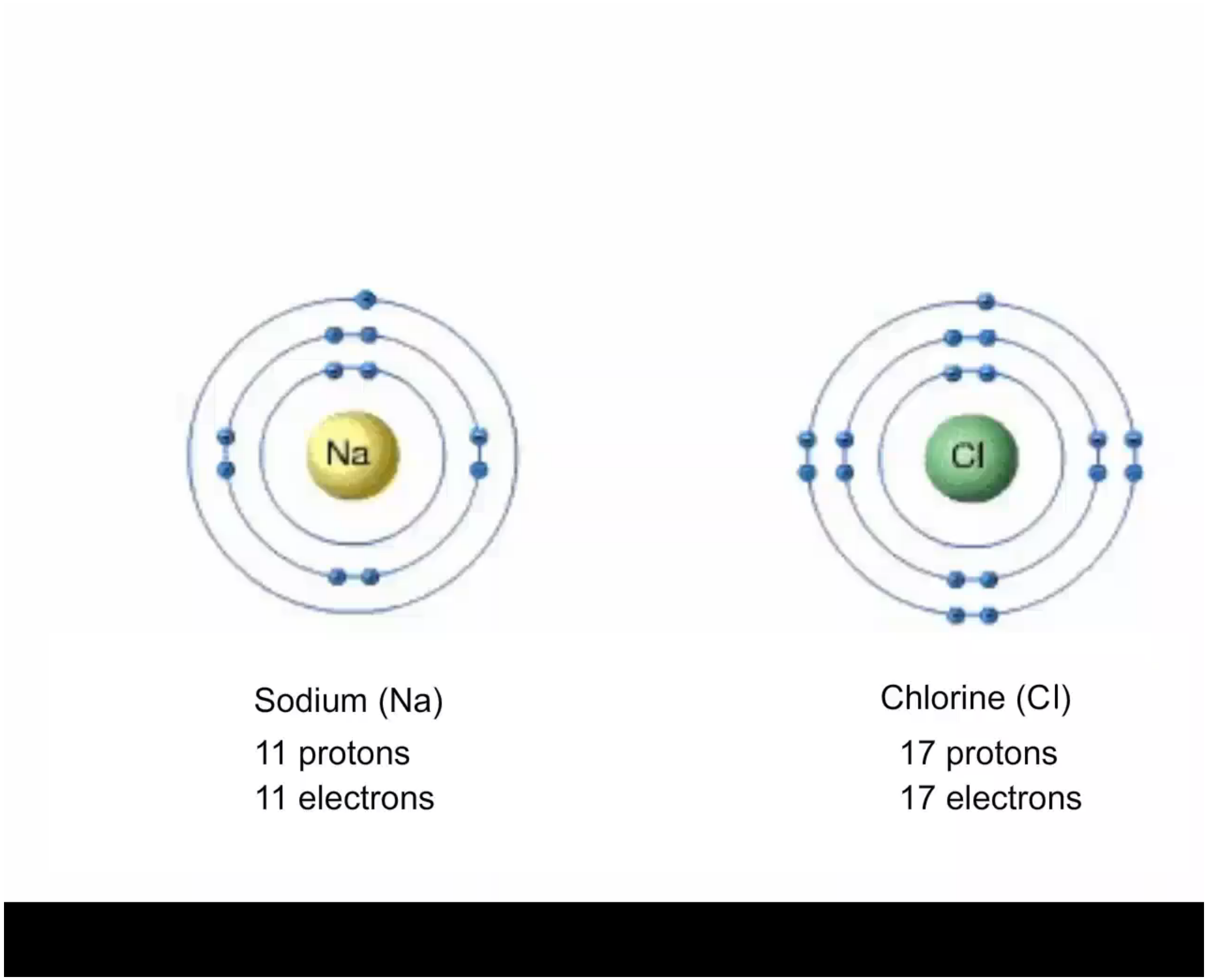
- Atoms sometimes strip electrons from their bonding partners
- The two resulting oppositely charged atoms or molecules are called **ions**
- A positively charged ion is called a **cation**
- A negatively charged ion is called an **anion**
- Anions and cations attract each other; this attraction is called an **ionic bond**

Figure 2.12



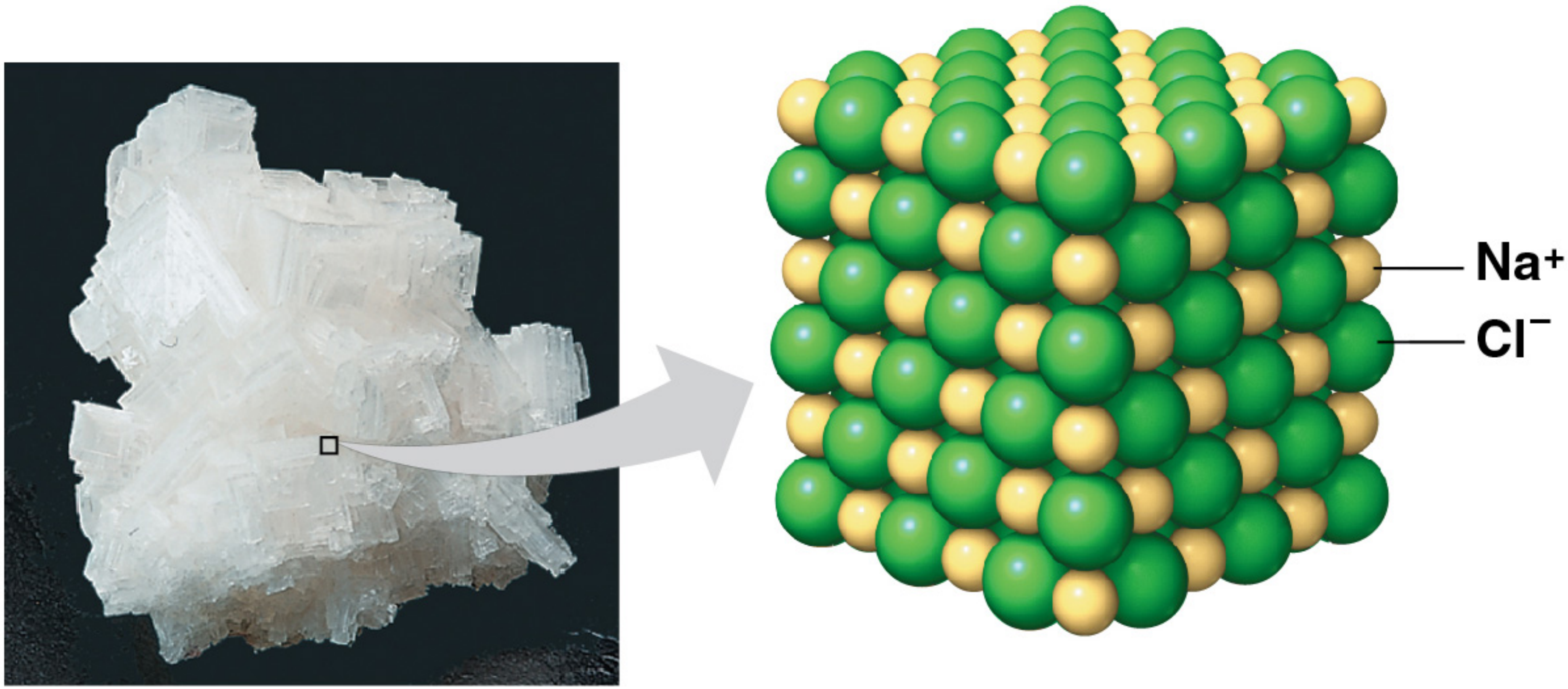
+
–
Na
Cl
Na
Cl
Na
Sodium atom
Cl
Chlorine atom
$Na^+$
Sodium ion
(a cation)
$Cl^-$
Chloride ion
(an anion)
Sodium chloride (NaCl)

# Animation: Ionic Bonds



- Compounds formed by ionic bonds are called **ionic compounds**, or **salts**
- Salts, such as sodium chloride (NaCl; table salt), are often found in nature as crystals
- NaCl itself is not a molecule; the formula for an ionic compound indicates the ratio of elements in a crystal of the salt
- Most salts are quite stable when dry, but dissociate quite easily in water

Figure 2.13



# Weak Chemical Interactions

- Most of the strongest bonds in organisms are covalent bonds that form a cell's molecules
- Many large biological molecules are held in their functional form by weak bonds
- The reversibility of weak bonds can be an advantage
- There are several types of weak chemical interactions that are important in organisms

## *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

# Animation: Hydrogen Bonds

**Hydrogen Bonds**

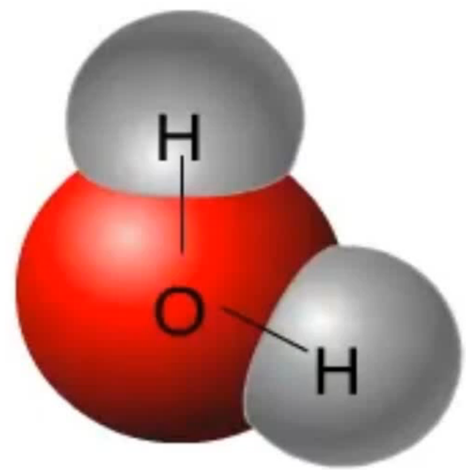
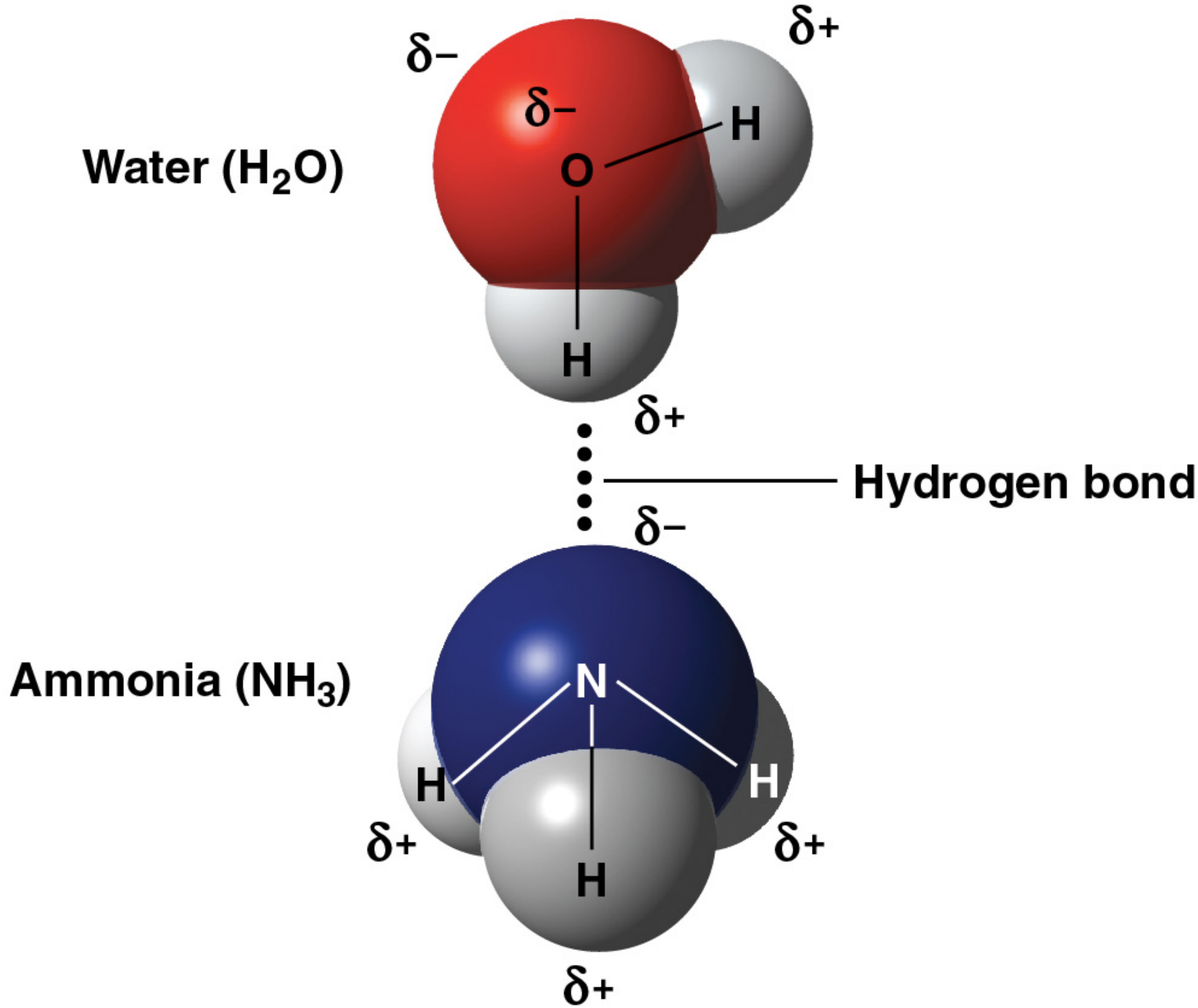


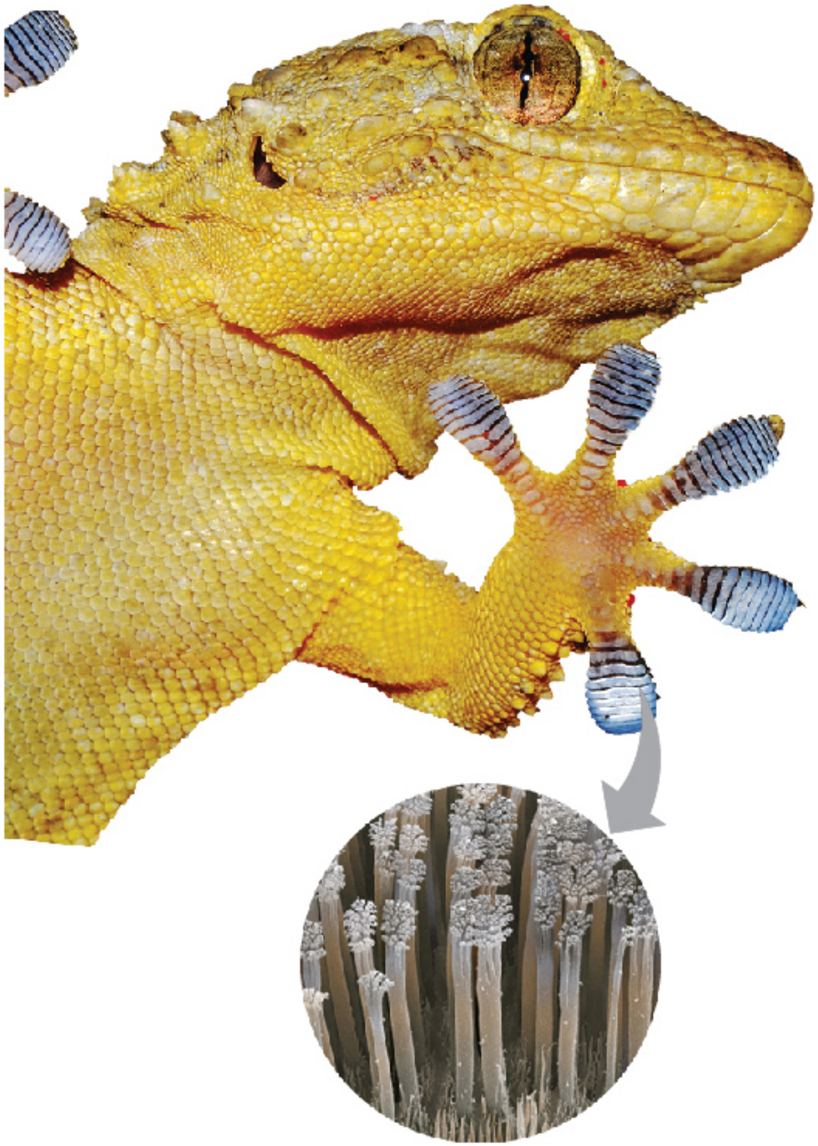
Figure 2.14



## *Van der Waals Interactions*

- If electrons are not evenly distributed, they may accumulate by chance in one part of a molecule
- **Van der Waals interactions** are attractions between molecules that are close together as a result of these charges
- Collectively, such interactions can be strong, as between molecules of a gecko’s toe hairs and a wall surface

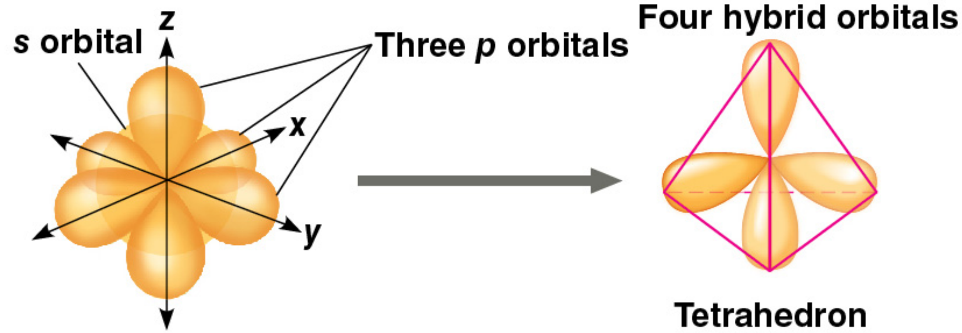
Figure 2.UN02



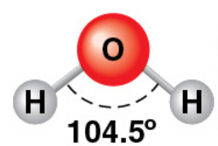
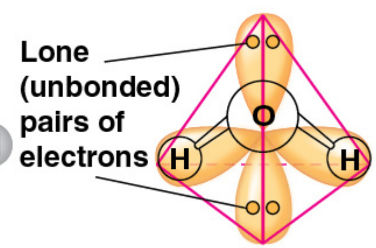
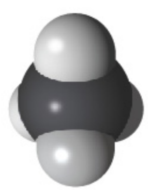
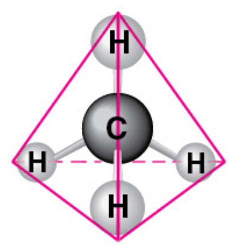
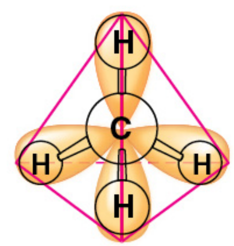
# Molecular Shape and Function

- A molecule’s size and shape are key to its function
- A molecule’s shape is determined by the positions of its atoms’ orbitals
- In a covalent bond, the *s* and *p* orbitals may hybridize, creating specific molecular shapes

Figure 2.15



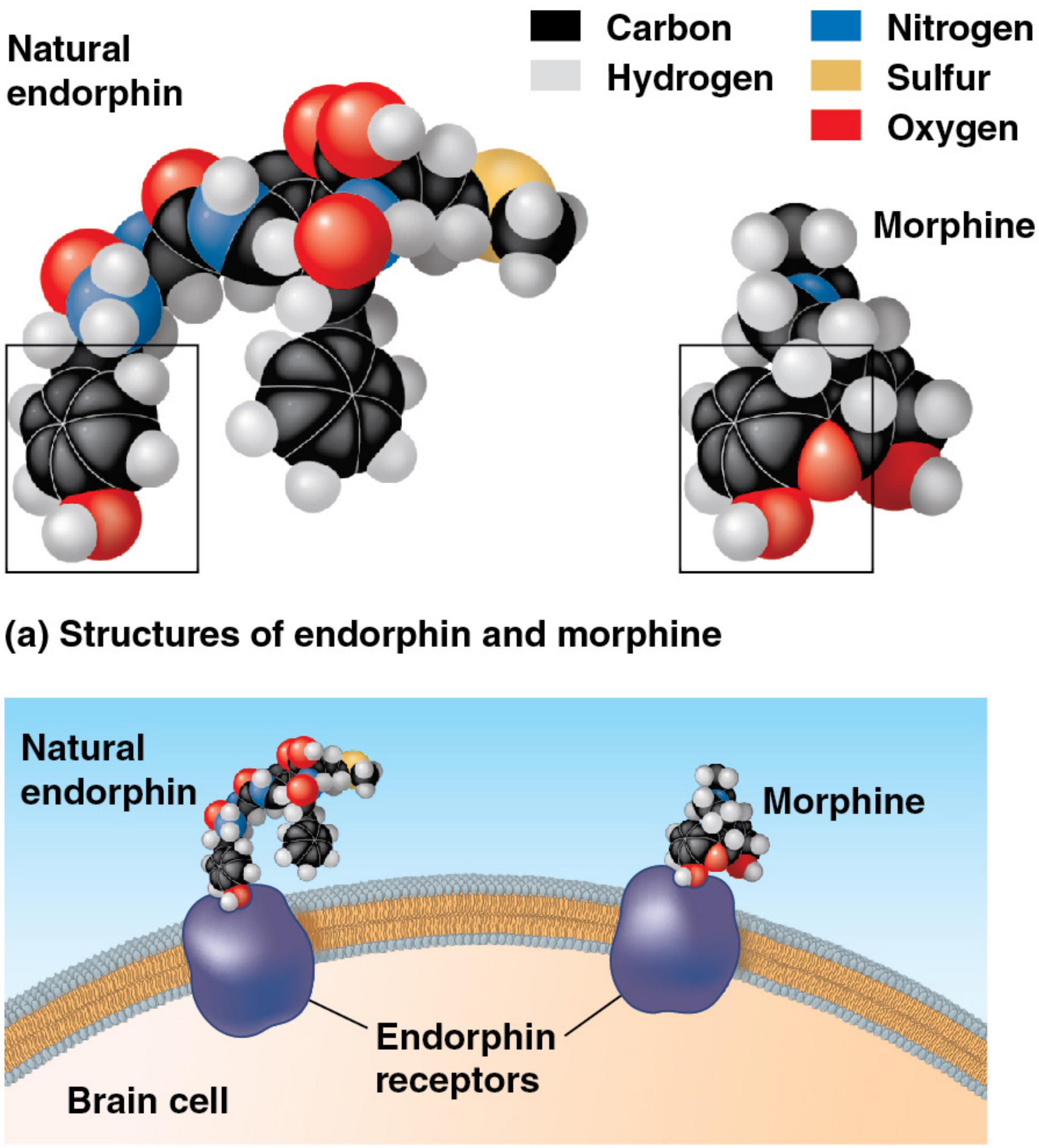

**(a) Hybridization of orbitals**

| Space-Filling Model | Ball-and-Stick Model | Hybrid-Orbital Model (with ball-and-stick model superimposed) |
|---|---|---|
|  Water ($H_2O$) |   |   |
|  Methane ($CH_4$) |   |   |

**(b) Molecular-shape models**

- Molecular shape determines how biological molecules recognize and respond to one another
- Opiates, such as morphine, and naturally produced endorphins have similar effects because their shapes are similar and they bind the same receptors in the brain

Figure 2.16

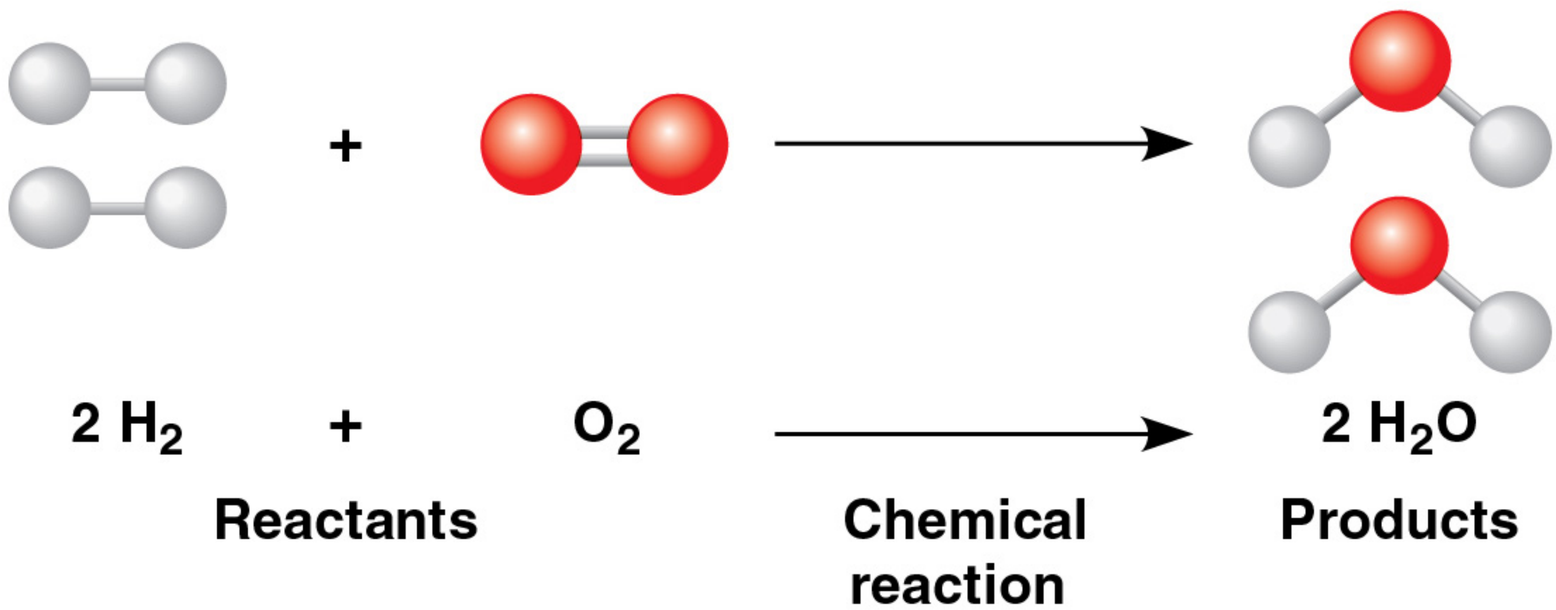



(b) Binding to endorphin receptors

# 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 resulting molecules of a chemical reaction are called **products**

Figure 2.UN03



- 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.UN04

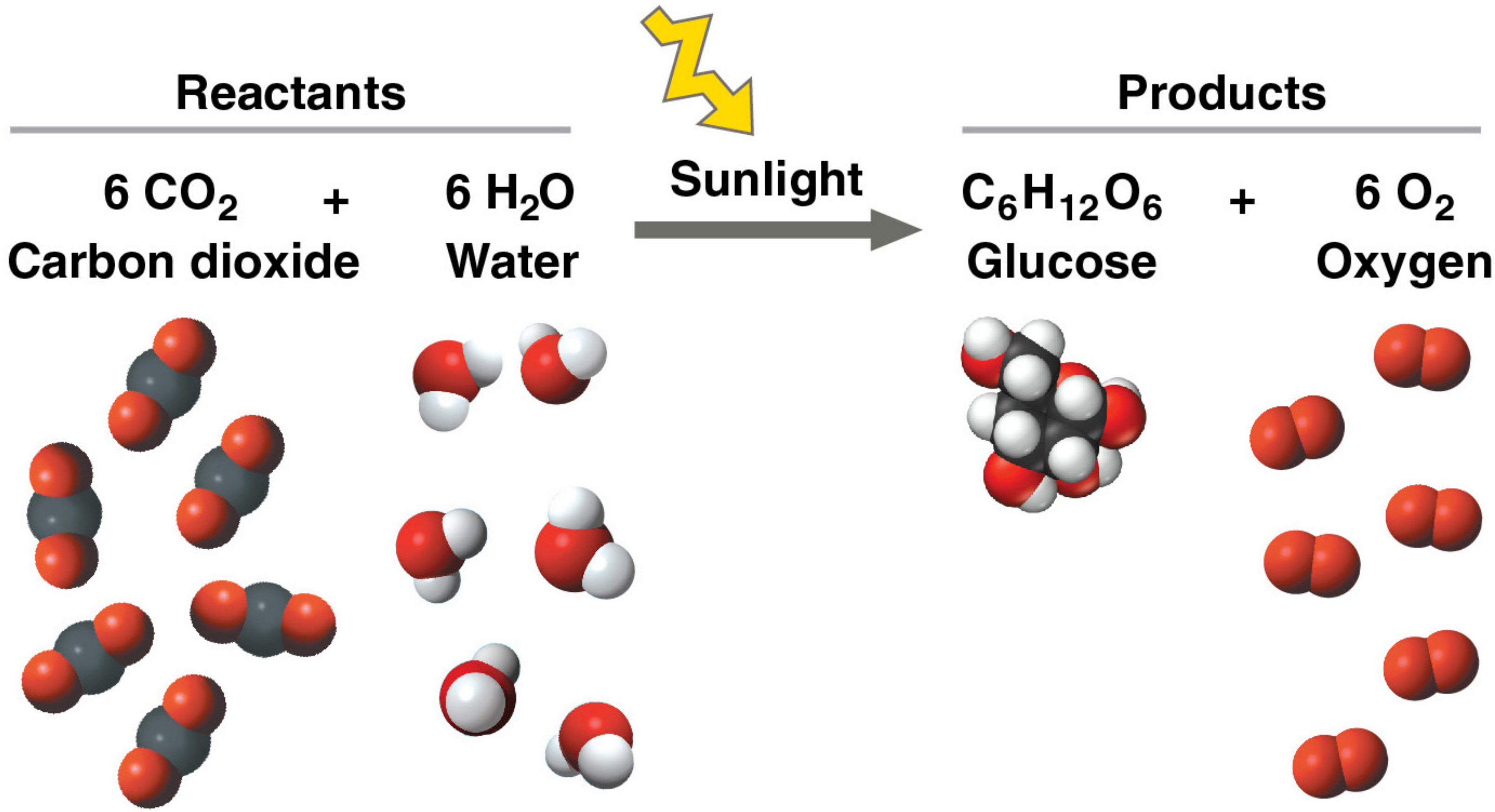


Figure 2.17

- All chemical reactions are reversible: Products of the forward reaction become reactants for the reverse reaction
- The two opposite-headed arrows indicate that a reaction is reversible

$$3\ H_2 + N \rightleftharpoons 2\ NH_3$$

- **Chemical equilibrium** is reached when the forward and reverse reactions occur at the same rate
- At equilibrium the relative concentrations of reactants and products do not change

Figure 2.UN05

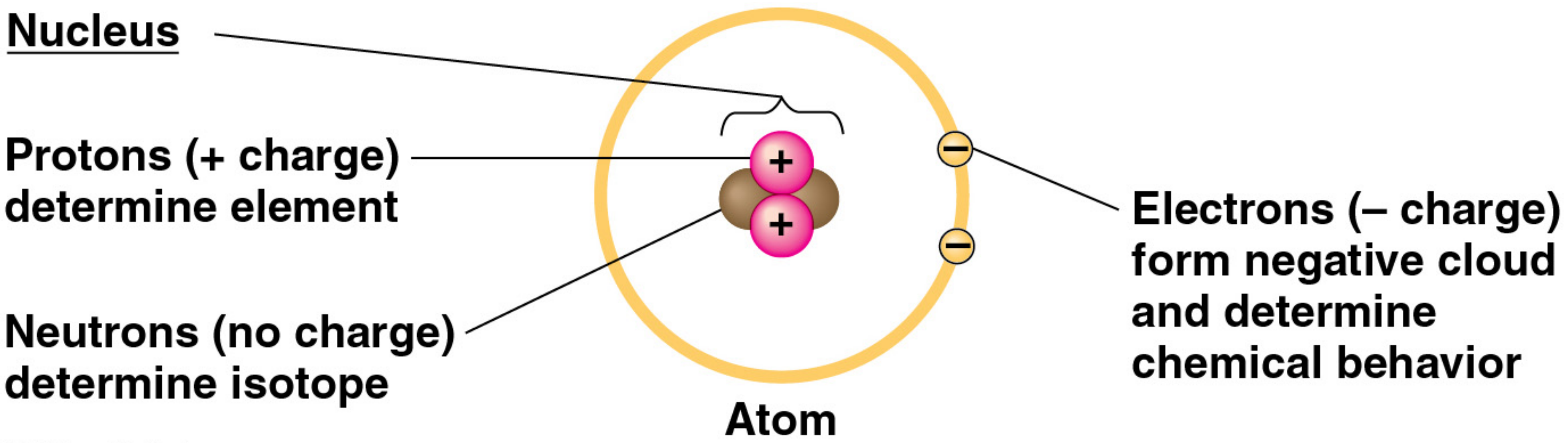


Figure 2.UN06

Figure 2.UN08

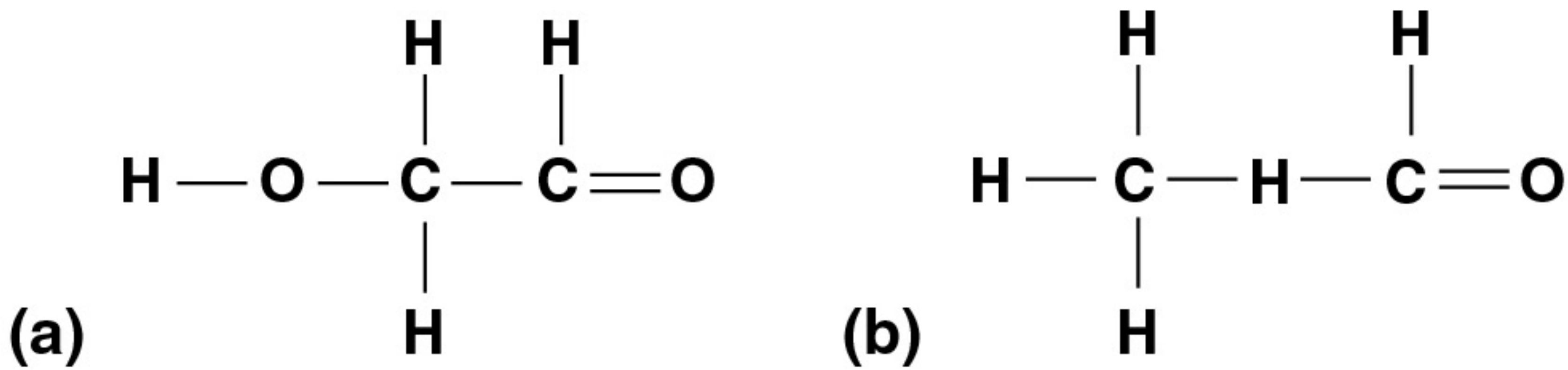


Figure 2.UN09

Figure 2.UN10