Chapter 2 – The Chemical Context of Life

I. Overview: Chemical Foundations of Biology

A. Living organisms and the world they live in are subject to the basic laws of physics and chemistry.
B. Biology is a multidisciplinary science, drawing on insights from other sciences.
C. Life can be organized into a hierarchy of structural levels.
D. At each successive level, additional emergent properties appear.

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

A. Organisms are composed of matter = is anything that takes up space and has mass, made up of elements

B. An **element** is a substance that cannot be broken down into other substances by chemical reactions.

1. 92 naturally occurring elements – rest are made in lab
2. Each element has a unique symbol, usually the first one or two letters of the name. Some of the symbols are derived from Latin or German names.

C. A **compound** is a substance consisting of two or more elements in a fixed ratio.

1. Table salt (sodium chloride or NaCl) is a compound with equal numbers of atoms of the elements chlorine and sodium.
2. While pure sodium (explosive metal) and chlorine (lethal gas), they combine to form edible and safe salt. This change in characteristics when elements combine to form a compound is an example of an emergent property.

D. 25 chemical elements are essential to life.

1. About 25 of the 92 natural elements are known to be **essential** for life.

a) Four elements—carbon (C), oxygen (O), hydrogen (H), and nitrogen (N)—make up 96% of living matter.

b) Trace elements are required by an organism but only in minute quantities.

*(1) Some **trace elements**, like iron (Fe), are required by all organisms.*

(2) Other trace elements are required by only some species.

Table 2.1 Naturally Occurring Elements in the Human Body

Symbol	Element	Atomic Number (See p. 29)	Percentage of Human Body Weight
O	Oxygen	8	65.0
C	Carbon	6	18.5
H	Hydrogen	1	9.5
N	Nitrogen	7	3.3
Ca	Calcium	20	1.5
P	Phosphorus	15	1.0
K	Potassium	19	0.4
S	Sulfur	16	0.3
Na	Sodium	11	0.2
Cl	Chlorine	17	0.2
Mg	Magnesium	12	0.1

Trace elements (less than 0.01%): 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), and zinc (Zn).

(a) daily intake of 0.15 milligrams of iodine is required for normal activity of the human thyroid gland.

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

A. Each element consists of unique **atoms**.

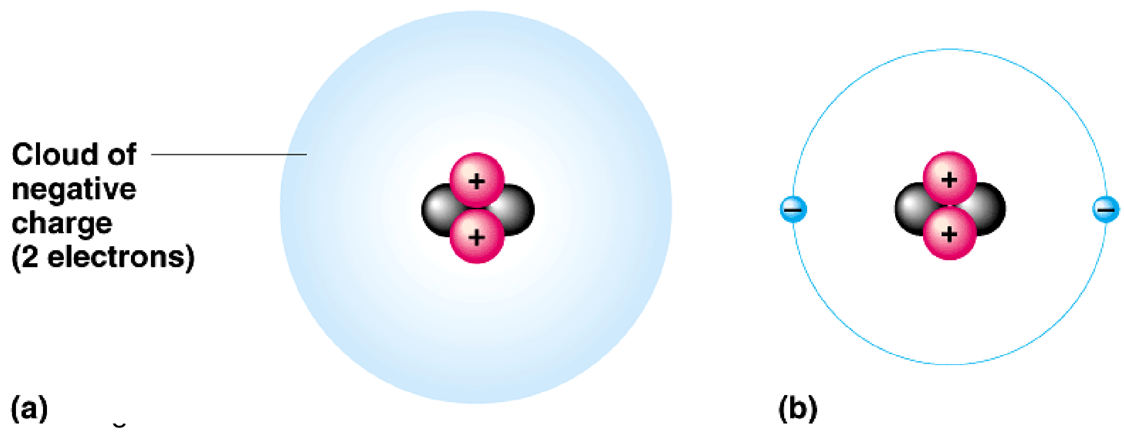
B. An atom is the smallest unit of matter that still retains the properties of an element.

C. Atoms are composed of even smaller parts, called **subatomic particles**.

1. **neutrons and protons**, are packed together to form a dense core, the atomic nucleus, at the center of an atom.

    a) Each proton has one unit of positive charge.

    b) N



electrically neutral.

2. **Electrons** orbit so fast, they can be visualized as forming a cloud of negative charge around the nucleus.

    a) Each electron has one unit of negative charge.

D. **Atomic Forces**: The attractions between opposite charges hold atom together

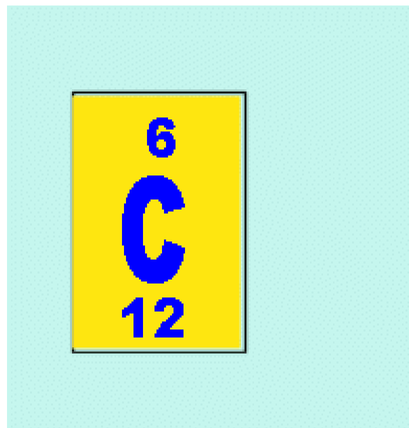
E. Protons have about the same mass as Neutrons (A neutron and a proton are almost identical in mass)

1. Someone came up with the Dalton - used to measure the mass of subatomic particles, atoms, or molecules - The mass of a neutron or a proton is close to 1 dalton
2. The mass of an electron is about 1/2000 that of a neutron or proton. - electrons too light to affect atomic mass

F. All atoms of a particular element have the same number of protons in their nuclei.

1. This number of protons is the element's unique **atomic number**.
2. The atomic number is written above the symbol for the element. For example, $_6C$ means that an atom of carbon has 6 protons in its nucleus.

G. Unless otherwise indicated, atoms have equal numbers of protons and electrons and, therefore, no **net charge.**



1. The atomic number tells us the number of protons and the number of electrons that are found in a neutral atom of a specific element.

H. The **mass number** is the sum of the number of protons and neutrons in the nucleus of an atom.

1. # of neutrons in an atom = the mass number - the number of protons (the atomic number)
2. symbol (for example, 12C).
3. The atomic weight of an atom, a measure of its mass, can be approximated by the mass number.
4. For example, 4He has a mass number of 4 and an estimated atomic weight of 4 daltons. (Actually, its atomic weight is 4.003 daltons.)
5. While all atoms of a given element have the same number of protons, they may differ in the number of neutrons.

I. Two atoms of the same element that differ in the number of neutrons are called **isotopes.**

1. In nature, an element occurs as a mixture of isotopes.

   a) For example, 99% of carbon atoms have 6 neutrons (12C). Most of the rest have 7 neutrons (13C) while the rarest carbon isotope, with 8 neutrons, is 14C.

2. Most isotopes are stable; they do not tend to lose particles. (Both ^{12}C and ^{13}C are **stable isotopes**.)

3. The nuclei of some isotopes are unstable and decay spontaneously, emitting particles and energy.

   a) ^{14}C is one of these **unstable isotopes**, or **radioactive isotopes**.

   b) When ^{14}C decays, one of its neutrons is converted to a proton and an electron.

   c) This converts ^{14}C to ^{14}N, transforming the atom to a different element.

4. Radioactive isotopes have many applications in biological research.

   a) Radioactive decay rates are predictable & can be used to date fossils.

   b) Radioactive isotopes can be used to trace atoms through metabolic processes. (feed an animal a nutrient w/ ^{14}C and you can follow where in the body that nutrient goes)

5. Radioactive isotopes are also used to diagnose medical disorders.

   a) For example, a known quantity of a substance labeled with a radioactive isotope can be injected into the blood, and its rate of excretion in the urine can be measured.

   b) Also, radioactive tracers can be used with imaging instruments to monitor chemical processes in the body.

6. While useful in research and medicine, the energy emitted in radioactive decay is hazardous to life.

   a) This energy can destroy molecules (usually DNA) within living cells.

   b) The severity of damage depends on the type and amount of radiation that the organism absorbs. (Sunburn to Nuclear Fallout)

J. Electron configuration influences the chemical behavior of an atom.

1. Atoms are mostly empty space.

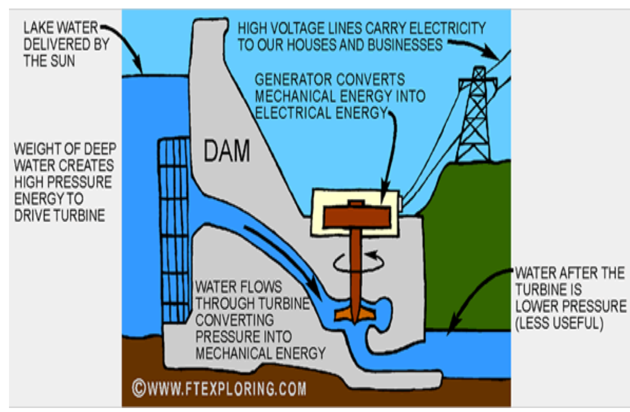
a) Atoms do <u>not</u> look like their cute diagrams. If the nucleus of this nitrogen atom was as big as it is drawn, the electrons would be about 150 meters from here (somewhere in the parking lot!).

b) When two elements interact during a chemical reaction, it is actually their electrons that are involved. The nuclei do not come close enough to interact.

K. The electrons of an atom vary in the amount of energy they possess.

1. **Energy** is the ability to do work.

2. **Potential energy** is the energy that matter stores because of its position or location. (because it doesn't like it's position or location)



a) Water stored behind a dam has potential energy that can be used to do work turning electric generators as it flows downhill.

b) Because potential energy has been expended, the water stores less energy at the bottom of the dam than it did in the reservoir.

3. Electrons have potential energy because of their position relative to the nucleus.

a) The negatively charged electrons are attracted to the positively charged nucleus.

b) The farther electrons are from the nucleus, the more potential energy they have. *(Rubber Band)*

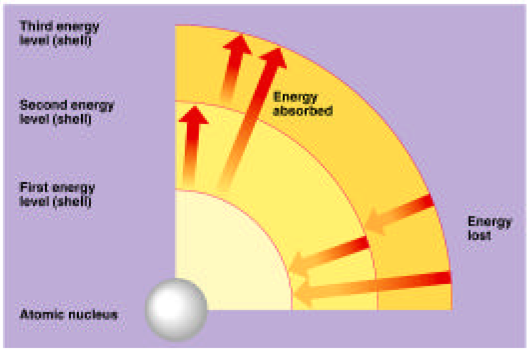
4. Electrons can only exist in certain fixed locations around the nucleus.

a) An electron cannot exist <u>between</u> these fixed locations. (Like you can't stand between two rungs of a ladder)

5. The different states of potential energy that the electrons of an atom can have are called **energy levels** or **electron shells**.

a) The first shell, closest to the nucleus, has the lowest potential energy. Electrons in outer shells have more potential energy.

b) Electrons can change their position only if they absorb or release a quantity of energy that matches the difference in potential energy between the two levels.



L. The chemical behavior of an atom is determined by its electron configuration—the distribution of electrons in its electron shells.

1. The first 18 elements, including those most important in biological processes, can be arranged in 8 columns and 3 rows.

2. Elements in the same row fill the same shells with electrons.

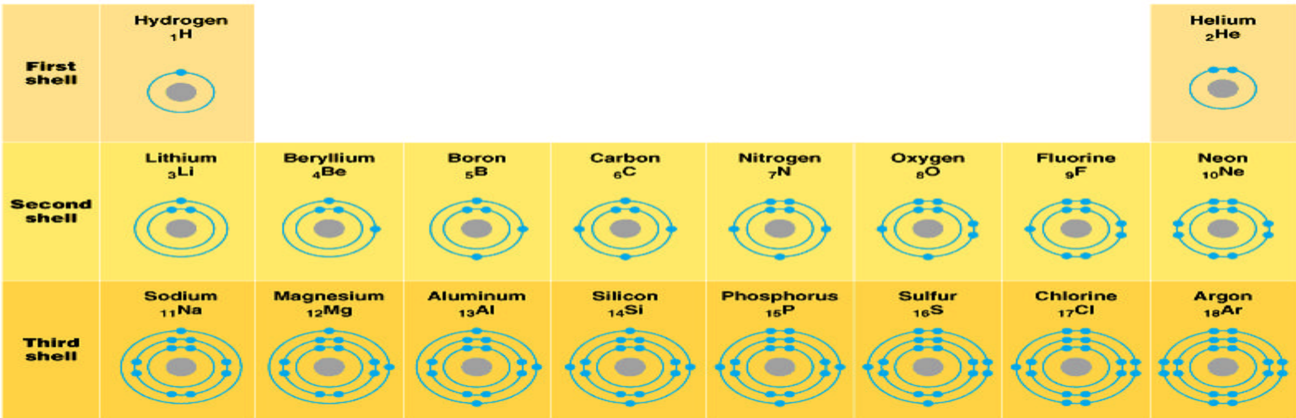
a) Moving from left to right, each element adds one electron (and proton) from the element before.

3. The first electron shell can <u>hold only 2</u> electrons.

a) The two electrons of helium fill the first shell.

4. Atoms with more than two electrons must place the extra electrons in higher shells.

a) All other shells can hold (and WANT to hold) <u>8 electrons.</u>



M. The chemical behavior of an atom depends mostly on the number of electrons in its outermost shell, the **valence shell**.

1. Electrons in the valence shell are known as valence electrons.
2. Lithium has one valence electron; neon has eight.

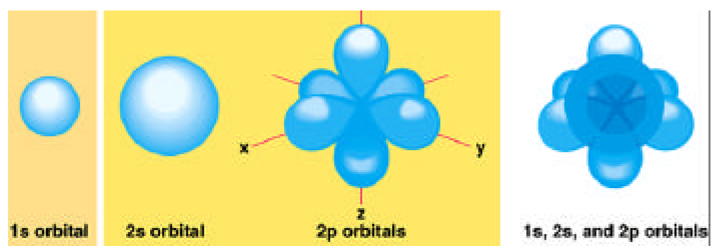
N. Atoms with the same number of **valence electrons** have similar chemical behaviors.

1. An atom with a **completed valence shell**, like neon or helium, is nonreactive. They do chemically react to form compounds with anything.

2. All other atoms are chemically reactive because they have **incomplete valence shells**.

a) Electrons orbit in pairs. The reactivity of an atom is because of unpaired electrons in one or more orbitals of their valence shells.
b) Gain more e^- to fill shell (C wants 4, N wants 3, O wants 2, H wants only 1)

3. The paths of electrons are often portrayed as concentric paths, like planets orbiting the sun.



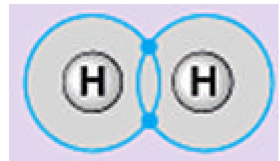
a) In reality, an e- occupies a more complex three-dimensional space, an **orbital**.

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

A. Atoms with incomplete valence shells can interact with each other by **sharing** or **transferring** valence electrons.

B. These interactions typically result in the atoms remaining close together, held by attractions called **chemical bonds**.

C. A **covalent bond** is formed by the sharing of a pair of valence electrons by two atoms. (Co = together / share: Cooperate)



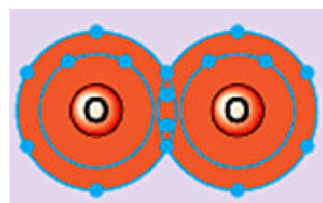
1. If two atoms come close enough that their unshared orbitals overlap, they will share their newly paired electrons. Each atom can count both electrons toward its goal of filling the valence shell.
2. two hydrogen atoms can share a pair of electrons, with each atom contributing one. This forms a **single covalent bond**

D. Two or more atoms held together by covalent bonds constitute a molecule.

1. We can abbreviate the structure of the molecule in several ways.

   a) H—H is the **structural formula** for the covalent bond between two hydrogen atoms.
   b) H_2 is the **molecular formula** for hydrogen gas.

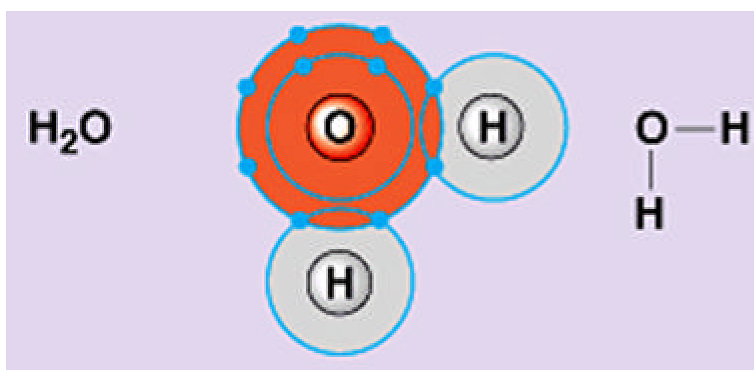
E. Oxygen needs to add 2 electrons to the 6 already present to complete its valence shell.



1. Two oxygen atoms can form a molecule by sharing *two* pairs of valence electrons.

2. These atoms have formed a **double covalent bond**.

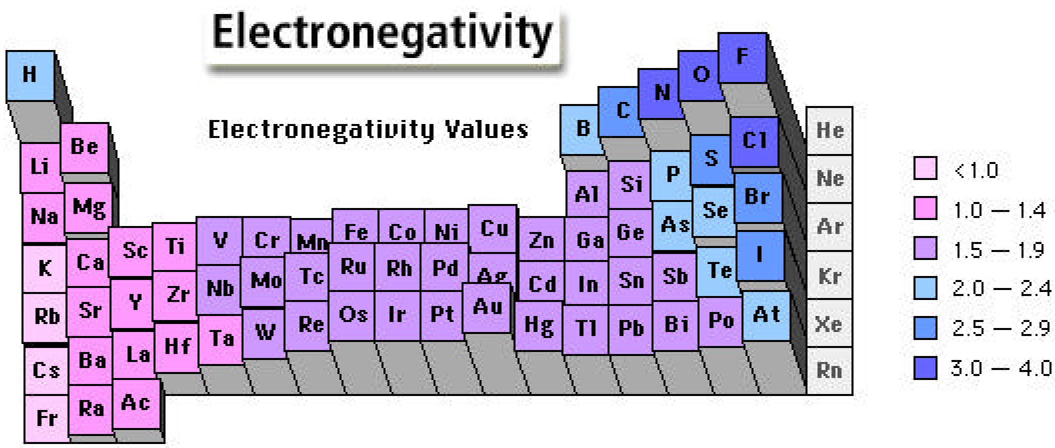
F. Covalent bonds can form between atoms of the same element or atoms of different elements.



1. While both types are molecules, the latter are also **compounds** (= more than one part)
2. Water, H_2O, is a compound in which two hydrogen atoms form single covalent bonds with an oxygen atom.

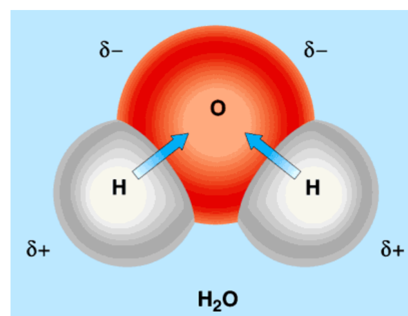
   a) This satisfies the valences of both elements.

G. The attraction of an atom for the shared electrons of a covalent bond is called its **electronegativity**.



1. Strongly electronegative atoms (Right of Per. Table) attempt to pull the shared electrons toward themselves. (The Greedy Take from the Generous)
2. Weakly electronegative atoms (Left of Per. Table) are not able to hold their electrons when faced with a bully.
3. Two Greedies (like O) have to share equally, two Generous (like H) will share equally. These are **non-polar covalent** bonds. The electrons (and their negative charges are equally distributed.

H. When two atoms that differ in electronegativity bond, they do not share the electron pair equally = a **polar covalent bond**.



1. The bonds between O and H in water are polar covalent because oxygen has a much higher electronegativity than does hydrogen.
2. Compounds with a polar covalent bond have regions of partial negative charge near the strongly electronegative atom and regions of partial positive charge near the weakly electronegative atom.

I. An **ionic bond** can form if two atoms are so unequal in their attraction for valence electrons that one atom (Bully) strips an electron completely from the other (Wimp).

1. For example, sodium, (1 valence electron), loses this electron to chlorine, (7 valence electrons).
2. Both atoms now have full valence shells, but are no longer neutral, but have charges and are called ions.

3. Sodium has one more proton than electrons and has a net positive charge. It is called a **Cation**

A

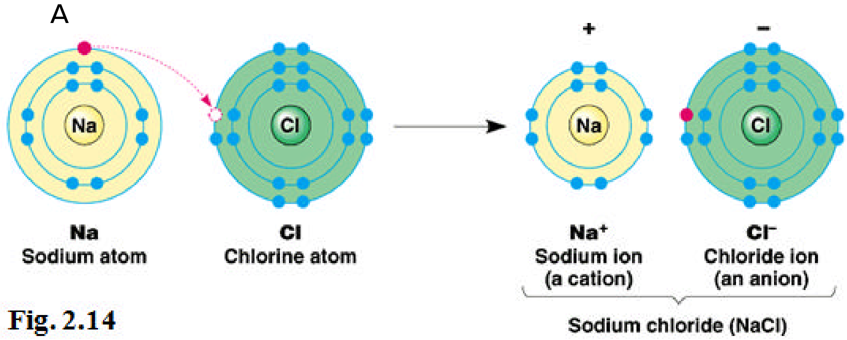


Fig. 2.14

4. Chlorine has one more electron than protons and has a net negative charge. It is called and **Anion**.

J. Because of differences in charge, cations and anions are attracted to each other to form an ionic bond.

1. Atoms in an ionic bond need not have acquired their charges by transferring electrons with each other.

K. Compounds formed by ionic bonds are **ionic compounds**, or **salts**. An example is NaCl, or table salt.

1. NaCl is not a molecule, but a salt crystal with equal numbers of Na^+ and Cl^- ions.

L. The strength of ionic bonds depends on environmental conditions, such as moisture.

1. Water can dissolve salts by reducing the attraction between the salt's anions and cations.

M. Weak chemical bonds play important roles in the chemistry of life.

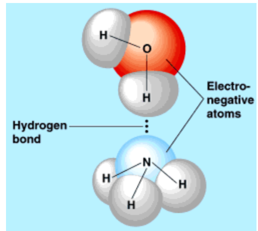
1. Within a cell, weak, brief bonds between molecules are important to a variety of processes.

a) For example, signal molecules from one neuron use weak bonds to bind briefly to receptor molecules on the surface of a receiving neuron. This triggers a response by the recipient.

N. Weak interactions include ionic bonds (weak in water), **hydrogen bonds**, and **van der Waals interactions**.

1. Hydrogen bonds form when a hydrogen atom already covalently bonded to a strongly electronegative atom is attracted to another strongly electronegative atom. (like 2 magnets)

a) These strongly electronegative atoms are typically nitrogen or oxygen.



b) The **partially positive**–charged hydrogen atom is attracted to regions of full or **partial negative** charge on molecules, atoms, or even regions of the same large molecule.

c) This explains some of waters weird properties. (What is this?)

2. Van der Waals forces can occur between any molecules.

a) Because electrons are constantly in motion, there can be periods when there happens to be more on one side of the molecule.

b) This creates ever-changing regions of partial negative and positive charge within a molecule.

c) Molecules or atoms in close proximity can be attracted by these fleeting charge differences, creating van der Waals interactions.

d) While individual bonds (ionic, hydrogen, van der Waals) are weak and temporary, collectively they are strong and play important biological roles.

A molecule's biological function is related to its shape.

The three-dimensional shape of a molecule is an important determinant of its function in a cell.
A molecule with two atoms is always linear.
However, a molecule with more than two atoms has a more complex shape.
The shape of a molecule is determined by the positions of the electron orbitals that are shared by the atoms involved in the bond.

- When covalent bonds form, the orbitals in the valence shell of each atom rearrange.

For atoms with electrons in both *s* and *p* orbitals, the formation of a covalent bonds leads to hybridization of the orbitals to four new orbitals in a tetrahedral shape.
In a water molecule, two of oxygen's four hybrid orbitals are shared with hydrogen atoms. The water molecule is shaped like a V, with its two covalent bonds spread apart at an angle of 104.5°.
In a methane molecule (CH_4), the carbon atom shares all four of its hybrid orbitals with H atoms. The carbon nucleus is at the center of the tetrahedron, with hydrogen nuclei at the four corners.
Large organic molecules contain many carbon atoms. In these molecules, the tetrahedral shape of carbon bonded to four other atoms is often a repeating motif.
Biological molecules recognize and interact with one another with a specificity based on molecular shape.
For example, signal molecules from a transmitting cell have specific shapes that bind to complementary receptor molecules on the surface of the receiving cell.

- The temporary attachment of the receptor and signal molecule stimulates activity in the receptor cell.

Molecules with similar shapes can have similar biological effects.

- For example, morphine, heroin, and other opiate drugs are similar enough in shape that they can bind to the same receptors as natural signal molecules called endorphins.
- Binding of endorphins to receptors on brain cells produces euphoria and relieves pain. Opiates mimic these natural endorphin effects.

Concept 2.4 Chemical reactions make and break chemical bonds

In chemical reactions, chemical bonds are broken and reformed, leading to new arrangements of atoms.
The starting molecules in the process are called reactants, and the final molecules are called products.
In a chemical reaction, all of the atoms in the reactants must be present in the products.

- The reactions must be "balanced".
- Matter is conserved in a chemical reaction.
- Chemical reactions rearrange matter; they do not create or destroy matter.

For example, we can recombine the covalent bonds of H_2 and O_2 to form the new bonds of H_2O.
In this reaction, two molecules of H_2 combine with one molecule of O_2 to form two molecules of H_2O.
Photosynthesis is an important chemical reaction.

- Humans and other animals ultimately depend on photosynthesis for food and oxygen.
- Green plants combine carbon dioxide (CO_2) from the air and water (H_2O) from the soil to create sugar molecules and release molecular oxygen (O_2)

as a by-product.

- This chemical reaction is powered by sunlight.
- The overall process of photosynthesis is $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$.
- This process occurs in a sequence of individual chemical reactions that rearrange the atoms of the reactants to form the products.

Some chemical reactions go to completion; that is, all the reactants are converted to products.
Most chemical reactions are reversible, with the products in the forward reaction becoming the reactants for the reverse reaction.
For example in this reaction: $3H_2 + N_2 \rightleftharpoons 2NH_3$ hydrogen and nitrogen molecules combine to form ammonia, but ammonia can decompose to hydrogen and nitrogen molecules.

- Initially, when reactant concentrations are high, they frequently collide to create products.
- As products accumulate, they collide to reform reactants.

Eventually, the rate of formation of products is the same as the rate of breakdown of products (formation of reactants), and the system is at chemical equilibrium.

- At equilibrium, products and reactants are continually being formed, but there is no net change in the concentrations of reactants and products.
- At equilibrium, the concentrations of reactants and products are typically not equal, but their concentrations have stabilized at a particular ratio.