Chapter 2

The Chemical Context of Life

Lecture Outline

Overview: A Chemical Connection to Biology

- Living organisms and the world they live in are subject to the basic laws of physics and chemistry.
- Biology is a multidisciplinary science, drawing on insights from other sciences.
- Life can be organized into a hierarchy of structural levels.
    - Atoms are organized into molecules, and molecules are organized into cells.
    - At each successive level, additional emergent properties appear.
    - Somewhere in the transition from molecules to cells, we cross the boundary between nonlife and life.

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

- Organisms are composed of **matter**, defined as anything that takes up space and has mass.
    - Matter is made up of elements.
- An **element** is a substance that cannot be broken down into other substances by chemical reactions.
    - There are 92 naturally occurring elements.
    - Each element has a unique symbol, usually the first one or two letters of its name. Some symbols are derived from Latin or German names.
- A **compound** is a substance that consists of two or more elements in a fixed ratio.
    - Table salt (sodium chloride or NaCl) is a compound with equal numbers of atoms of the elements chlorine and sodium.
    - Although pure sodium is a metal and chlorine is a gas, they combine to form an edible compound.
    - This change in characteristics when elements combine to form a compound is an example of an emergent property.

Twenty-five chemical elements are essential for life.

- About 20–25% of the 92 naturally occurring elements are **essential elements** that are necessary for living things.

- There is some variation: humans have 25 essential elements, while plants have only 17.

- Four elements—oxygen (O), carbon (C), hydrogen (H), and nitrogen (N)—make up 96% of living matter.
  - Most of the remaining 4% of an organism's weight consists of calcium (Ca), phosphorus (P), potassium (K), and sulfur (S).
- **Trace elements** are required by an organism but only in minute quantities.
  - Some trace elements, like iron (Fe), are required by all organisms.
  - Other trace elements are required by only some species.
    - For example, vertebrates (animals with backbones) require the element iodine (I) for normal activity of the thyroid gland. A daily intake of only 0.15 milligram (mg) of iodine is adequate for normal activity of the human thyroid.
- Some naturally occurring elements, such as arsenic, are toxic to most organisms.
  - Some species have become adapted to environments containing toxic elements.
  - For example, serpentine plant communities are able to survive in soils containing chromium, nickel, and cobalt.

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.
- Atoms are composed of even smaller parts called *subatomic particles*.
  - Two subatomic particles, **neutrons** and **protons**, are packed together to form a dense core, the **atomic nucleus**, at the center of an atom.
  - **Electrons** can be visualized as forming a cloud of negative charge around the nucleus.
- Each electron has one unit of negative charge, each proton has one unit of positive charge, and neutrons are electrically neutral.
  - The attractions between the positive charges in the nucleus and the negative charges of the electrons keep the electrons in the vicinity of the nucleus.
- A neutron and a proton have almost identical mass: about 1.7×10^{-24} gram per particle.
- For convenience, a smaller unit, the **dalton,** is used to measure the masses of subatomic particles, atoms, and molecules.
  - The mass of a neutron or a proton is close to 1 dalton.
- The mass of an electron is about 1/2,000 the mass of a neutron or proton.
  - We typically ignore the contribution of electrons when determining the total mass of an atom.
- All atoms of a particular element have the same number of protons in their nuclei.
  - This number of protons is the element's unique **atomic number**.
  - The atomic number is written as a subscript before the symbol for the element. For example, $_2He$ means that an atom of helium has 2 protons in its nucleus.

- Unless otherwise indicated, atoms have equal numbers of protons and electrons and, therefore, no net charge.
    - Therefore, the atomic number tells the number of protons and the number of electrons in a neutral atom of a specific element.
- The **mass number** is the sum of the number of protons and the number of neutrons in the nucleus of an atom.
    - We can determine the number of neutrons in an atom by subtracting the number of protons (the atomic number) from the mass number.
    - The mass number is written as a superscript before an element's symbol (for example, ^{4}He).
    - Because neutrons and protons each have a mass very close to 1 dalton, the mass number is an approximation of the total mass of an atom, called its **atomic mass.**
- Two atoms of the same element that differ in the number of neutrons are called **isotopes.**
- In nature, an element occurs as a mixture of isotopes.
    - For example, 99% of carbon atoms have 6 neutrons (^{12}C).
    - Most of the remaining 1% of carbon atoms have 7 neutrons (^{13}C), while the rarest carbon isotope, with 8 neutrons, is ^{14}C.
- Most isotopes are stable; they do not tend to lose particles.
    - Both ^{12}C and ^{13}C are stable isotopes.
- The nuclei of some isotopes are unstable and decay spontaneously, emitting particles and energy.
    - ^{14}C is one of these unstable isotopes, or **radioactive isotopes**.
    - When ^{14}C decays, one of its neutrons is converted to a proton and an electron. In this process, ^{14}C is converted to ^{14}N, a different element.
- Radioactive isotopes have many applications in biological research.
    - Radioactive decay rates can be used to date fossils.
    - Radioactive isotopes can be used to trace atoms through metabolic processes.
- Radioactive isotopes are also used to diagnose medical disorders.
    - 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.
    - Radioactive tracers can be used with imaging instruments such as PET scanners to monitor the body's chemical processes, such as those involved in cancerous growths.
- Though useful in research and medicine, the energy emitted in radioactive decay is hazardous to life.
    - This energy can damage molecules within living cells.
    - The severity of damage depends on the type and amount of radiation that the organism absorbs.

Electron configuration influences the chemical behavior of an atom.

- Simplified models of the atom greatly distort the atom's relative dimensions.
- In an accurate representation of the relative proportions of an atom, if an atom of helium were the size of Yankee Stadium, the nucleus would be only the size of a pencil eraser in the center of the field.

- The electrons would be like two gnats buzzing around the stadium.
- Atoms are mostly empty space.

- 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.
- The electrons of an atom vary in the amounts of energy they possess.
- **Energy** is the ability to cause change, by doing work.
- **Potential energy** is the energy that matter stores because of its structure or location.
  - Water stored behind a dam has potential energy due to its altitude.
  - That potential energy can be used to do work turning electric generators.
  - Because potential energy has been expended, the water stores less energy at the bottom of the dam than it did in the reservoir.
- Electrons have potential energy because of their positions relative to the nucleus.
  - The negatively charged electrons are attracted to the positively charged nucleus.
  - The farther electrons are from the nucleus, the more potential energy they have.
- Changes in an electron's potential energy can occur only in steps of a fixed amount, moving the electron to a fixed location relative to the nucleus.
  - An electron cannot exist between these fixed locations.
- The different states of potential energy of the electrons of an atom are called **electron shells.**
  - The first shell, closest to the nucleus, has the lowest potential energy.
  - Electrons in outer shells have higher potential energy.
  - 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.
- The chemical behavior of an atom is determined by its electron configuration—the distribution of electrons in its electron shells.
- An abbreviated *periodic table of the elements* shows the distribution of electrons in the first 18 elements from hydrogen to argon.
  - The elements are arranged in three rows or periods, corresponding to the number of electron shells in their atoms.
  - Elements in the same row have the same shells filled with electrons.
  - As we move from left to right in the table, each element has one more electron (and proton) than the element before.
- The first electron shell can hold only 2 electrons.
  - The 2 electrons of helium fill the first shell.
- Atoms with more than 2 electrons must have the extra electrons in higher shells.
  - Lithium, with 3 electrons, has 2 electrons in the first shell and 1 in the second shell.
- The second shell can hold up to 8 electrons.
  - Neon, with 10 total electrons, has 2 in the first shell and 8 in the second, thus filling both shells.

- The chemical behavior of an atom depends mostly on the number of electrons in its outermost shell, the **valence shell.**
    - Electrons in the valence shell are known as **valence electrons.**
    - Lithium has 1 valence electron; neon has 8.
- Atoms with the same number of valence electrons have similar chemical behaviors.
- An atom with a completed valence shell, like neon, is *inert* or chemically nonreactive.
- All other atoms are chemically reactive because they have incomplete valence shells.
- The paths of electrons are often portrayed as concentric paths, like planets orbiting the sun. In reality, an electron occupies a more complex three-dimensional space, an **orbital.**
- The orbital is the space in which the electron is found 90% of the time.
    - Each orbital can hold a maximum of 2 electrons.
    - The first shell has room for a single spherical 1*s* orbital for its pair of electrons.
    - The second shell can pack pairs of electrons into one spherical 2*s* orbital and three dumbbell-shaped 2*p* orbitals.
- The reactivity of atoms arises from the presence of unpaired electrons in one or more orbitals of their valence shells.
    - Electrons occupy separate orbitals within the valence shell until forced to share orbitals.
    - The 4 valence electrons of carbon each occupy separate orbitals, but the 5 valence electrons of nitrogen are distributed into three unshared orbitals and one shared orbital.
- When atoms interact to complete their valence shells, it is the *unpaired* electrons that are involved.

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

- Atoms with incomplete valence shells can interact with each other by sharing or transferring valence electrons.
- These interactions typically result in the atoms remaining close together, held by attractions called **chemical bonds.**
- The strongest chemical bonds are covalent bonds and ionic bonds.
- A **covalent bond** is formed when two atoms share a pair of valence electrons.
    - If two atoms come close enough that their unshared orbitals overlap, they can share their newly paired electrons. Each atom can count both electrons toward its goal of filling the valence shell.
    - For example, if two hydrogen atoms come close enough that their 1*s* orbitals overlap, then they can share a pair of electrons, with each atom contributing one.
- Two or more atoms held together by covalent bonds constitute a **molecule**.
- The *molecular formula* of a molecule indicates that the molecule consists of two atoms of hydrogen.
    - H_2 is the molecular formula for hydrogen gas.
- The *structural formula* uses a line for a **single bond**, a pair of shared electrons.

- H—H is the structural formula for the covalent bond between two hydrogen atoms.

- Electron sharing can be shown by an electron-distribution diagram or *Lewis dot structure*, in which element symbols are surrounded by dots representing valence electrons (H:H).
- Oxygen needs to add 2 electrons to the 6 already present to complete its valence shell.
  - Two oxygen atoms can form a molecule by sharing *two* pairs of valence electrons. These atoms form a **double covalent bond** (O=O).
- Every atom has a characteristic total number of covalent bonds that it can form, equal to the number of unpaired electrons in the outermost shell. This bonding capacity is called the atom's **valence.**
  - The valence of hydrogen is 1; oxygen is 2; nitrogen is 3; carbon is 4.
  - The situation is more complex for elements in the third row of the periodic table.
  - Phosphorus should have a valence of 3, based on its 3 unpaired electrons, but in biological molecules, it generally has a valence of 5, forming three single covalent bonds and one double bond.
- Covalent bonds can form between atoms of the same element (forming pure elements) or atoms of *different* elements (forming compounds).
  - Water (H_2O) is a compound in which two hydrogen atoms form single covalent bonds with an oxygen atom. This satisfies the valences of both elements.
  - Methane (CH_4) is a compound that satisfies the valences of both C and H.
- Atoms differ in the degree of their attraction for shared electrons. The attraction of an atom for the shared electrons of a covalent bond is called its **electronegativity**.
  - Strongly electronegative atoms strongly pull the shared electrons toward themselves.
- Electrons in a covalent bond are shared equally in a **nonpolar covalent bond**.
  - A covalent bond between two atoms of the same element is always nonpolar.
  - A covalent bond between atoms that have similar electronegativities is also nonpolar.
  - Because carbon and hydrogen do not differ greatly in electronegativities, the bonds of CH_4 are nonpolar.
- When two atoms that differ in electronegativity bond, they do not share the electron pair equally, and they form a **polar covalent bond**.
  - When oxygen and hydrogen form water, the bonds between oxygen and hydrogen are polar covalent because oxygen has a much higher electronegativity than does hydrogen.
  - Compounds with a polar covalent bond have regions of partial negative charge ($\delta-$) near the strongly electronegative atom and regions of partial positive charge ($\delta+$) near the weakly electronegative atom.
- An **ionic bond** can form if two atoms are so unequal in their attraction for valence electrons that one atom strips an electron completely from the other.
  - For example, sodium, with 1 valence electron in its third shell, transfers this electron to chlorine, with 7 valence electrons in its third shell.
  - Then sodium has a full valence shell (the second) and chlorine has a full valence shell (the third).
- After the transfer, both atoms are no longer neutral but have charges and are called **ions**.

- Atoms with positive charges are **cations**.
  - Sodium has one more proton than electrons and has a net positive charge (1+).
- Atoms with negative charges are **anions**.
  - Chlorine has one more electron than protons and has a net negative charge (1–).
- Because of differences in charge, cations and anions are attracted to each other to form an **ionic bond**.
  - Atoms in an ionic bond need not have acquired their charges by transferring electrons with each other.
- Compounds formed by ionic bonds are **ionic compounds**, or **salts**. An example is NaCl, or table salt.
  - The formula for an ionic compound indicates the ratio of elements in a crystal of that salt. NaCl is not a molecule but a salt crystal with equal numbers of Na^+ and Cl^- ions.
- Ionic compounds can have ratios of elements different from 1:1.
  - For example, the ionic compound magnesium chloride ($MgCl_2$) has two chloride atoms for each magnesium atom.
  - Magnesium needs to lose 2 electrons to drop to a full outer shell; each chlorine atom needs to gain 1 electron.
- Entire molecules that have full electrical charges are also called *ions*.
  - In the salt ammonium chloride (NH_4Cl), the anion is chloride (Cl^-) and the cation is ammonium (NH_4^+).
- The strength of ionic bonds depends on environmental conditions, such as moisture.
- Water can dissolve salts by reducing the attraction between the salt's anions and cations.

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

- Within a cell, weak, brief bonds between molecules are important to a variety of processes.
  - Many large biological molecules are held in their functional form by weak bonds.
  - When two molecules in the cell make contact, they may adhere temporarily by weak bonds.
- The reversibility of weak bonding can be an advantage: Two molecules can come together, respond to each other in some way, and then separate.
- Weak interactions include ionic bonds between ions dissociated in water, hydrogen bonds, and van der Waals interactions.
- **Hydrogen bonds** form when a hydrogen atom that is already covalently bonded to one electronegative atom is attracted to another electronegative atom.
  - In cells, the electronegative partners are typically nitrogen or oxygen.
  - Hydrogen bonds form because a polar covalent bond leaves the hydrogen atom with a partial positive charge and the other atom with a partial negative charge.
  - 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.
- For example, ammonia molecules and water molecules interact with weak hydrogen bonds.

- In the ammonia molecule, the hydrogen atoms have partial positive charges, and the more electronegative nitrogen atom has a partial negative charge.
- In the water molecule, the hydrogen atoms have partial positive charges, and the oxygen atom has a partial negative charge.
- Areas with opposite charges are attracted to each other.

- Even molecules with nonpolar covalent bonds can have temporary regions of partial negative and positive charge.
  - Because electrons are constantly in motion, there may be times when they accumulate by chance in one area of a molecule.
  - This creates ever-changing regions of partial negative and positive charge within a molecule.
- Molecules or atoms in close proximity can be attracted by fleeting charge differences, creating **van der Waals interactions**.
  - The resulting ever-changing regions of positive and negative charge enable atoms and molecules to stick to one another.
  - van der Waals interactions between molecules on the tips of hundreds of thousands of tiny hairs on a gecko toe and the molecules of the wall's surface are so numerous that, despite their individual weakness, together they can support the gecko's body weight, allowing the animal to walk up a wall.
- Although individual bonds (ionic, hydrogen, van der Waals) are temporary and individually weak, collectively they are strong and play important biological roles.
  - The cumulative effect of individual bonds is to reinforce the three-dimensional shape of a large molecule.

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, but most molecules with more than two atoms have 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 bond 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.

- 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.
    - The 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, the covalent bonds of H_2 and O_2 can recombine to form the new bonds of H_2O. 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 they release molecular oxygen (O_2) as a by-product. This chemical reaction is powered by sunlight.
    - The overall process of photosynthesis is $6\ CO_2 + 6\ H_2O \rightarrow C_6H_{12}O_6 + 6\ O_2$.
    - This process occurs in a sequence of individual chemical reactions that rearrange the atoms of the reactants to form the products.
- All chemical reactions are reversible, with the products in the forward reaction becoming the reactants in the reverse reaction.
    - For example, in the reaction $3\ H_2 + N_2 \leftrightarrow 2\ NH_3$, hydrogen and nitrogen molecules combine to form ammonia, but ammonia can decompose to hydrogen and nitrogen molecules.
    - Initially, when reactant concentrations are high, the reactants 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.
- Some chemical reactions go to completion; that is, all the reactants are converted to products.