Chapter 2: THE CHEMICAL CONTEXT OF LIFE

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What is Chemistry? Why Do We Need to Study Some Chemistry Within Biology?

• Chemistry is the science dealing with the **composition and properties of substances (physical matter)**, and with the **reactions** by which substances are produced from or converted into other substances.

• **Matter** is anything that takes up space and has mass. It consists of **chemical elements**, substances that cannot be broken down into simpler substances.

• Humans and all other living things are made of multiple chemicals. To study and understand **how living things function and interact with the environment**, we must have a basic understanding of chemistry.
How Matter Is Organized: Chemical Elements

- **Element**: Any substance that cannot be broken down to any other substance by chemical reactions; an element is made up entirely of one type of atom.

- Chemical elements:
  - Are the building blocks for living and nonliving matter.
  - Have specific properties that distinguish each element; these properties depend on the structure of the element’s atoms.
  - Are designated by a chemical symbol, usually the first letter or two of the element’s name (in English, Latin, or German).
Life Has a Unique Chemistry

- Chemists recognize 92 elements occurring in nature. About 25 are known to be essential to life, with 4 predominant elements:
  - oxygen \((O)\), carbon \((C)\), hydrogen \((H)\), and nitrogen \((N)\) make up 96% of living matter.

- Other important elements in living things that account for the remaining 4% of an organism’s weight include:
  - calcium \((Ca)\), phosphorus \((P)\), potassium \((K)\), sulfur \((S)\), sodium \((Na)\), chlorine \((Cl)\), magnesium \((Mg)\), and a few more trace elements (required in minute quantities).
### TABLE 2.1 Main Chemical Elements in the Body

<table>
<thead>
<tr>
<th>Chemical Element (Symbol)</th>
<th>% Of Total Body Mass</th>
<th>Significance</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>MAJOR ELEMENTS</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Oxygen (O)</td>
<td>65.0</td>
<td>Part of water and many organic (carbon-containing) molecules; used to generate ATP, a molecule used by cells to temporarily store chemical energy.</td>
</tr>
<tr>
<td>Carbon (C)</td>
<td>18.5</td>
<td>Forms backbone chains and rings of all organic molecules: carbohydrates, lipids (fats), proteins, and nucleic acids (DNA and RNA).</td>
</tr>
<tr>
<td>Hydrogen (H)</td>
<td>9.5</td>
<td>Constituent of water and most organic molecules; ionized form (H(^+)) makes body fluids more acidic.</td>
</tr>
<tr>
<td>Nitrogen (N)</td>
<td>3.2</td>
<td>Component of all proteins and nucleic acids.</td>
</tr>
<tr>
<td><strong>LESSER ELEMENTS</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Calcium (Ca)</td>
<td>1.5</td>
<td>Contributes to hardness of bones and teeth; ionized form (Ca(^{2+})) needed for blood clotting, release of some hormones, contraction of muscle, and many other processes.</td>
</tr>
<tr>
<td>Phosphorus (P)</td>
<td>1.0</td>
<td>Component of nucleic acids and ATP; required for normal bone and tooth structure.</td>
</tr>
<tr>
<td>Potassium (K)</td>
<td>0.35</td>
<td>Ionized form (K(^+)) is the most plentiful cation (positively charged particle) in intracellular fluid; needed to generate action potentials.</td>
</tr>
<tr>
<td>Sulfur (S)</td>
<td>0.25</td>
<td>Component of some vitamins and many proteins.</td>
</tr>
<tr>
<td>Sodium (Na)</td>
<td>0.2</td>
<td>Ionized form (Na(^+)) is the most plentiful cation in extracellular fluid; essential for maintaining water balance; needed to generate action potentials.</td>
</tr>
<tr>
<td>Chlorine (Cl)</td>
<td>0.2</td>
<td>Ionized form (Cl(^-)) is the most plentiful anion (negatively charged particle) in extracellular fluid; essential for maintaining water balance.</td>
</tr>
<tr>
<td>Magnesium (Mg)</td>
<td>0.1</td>
<td>Ionized form (Mg(^{2+})) needed for action of many enzymes, molecules that increase the rate of chemical reactions in organisms.</td>
</tr>
<tr>
<td>Iron (Fe)</td>
<td>0.005</td>
<td>Ionized forms (Fe(^{2+}) and Fe(^{3+})) are part of hemoglobin (oxygen-carrying protein in red blood cells) and some enzymes (proteins that catalyze chemical reactions in living cells).</td>
</tr>
<tr>
<td><strong>TRACE ELEMENTS</strong></td>
<td>0.2</td>
<td>Aluminum (Al), Boron (B), Chromium (Cr), Cobalt (Co), Copper (Cu), Fluorine (F), Iodine (I), Manganese (Mn), Molybdenum (Mo), Selenium (Se), Silicon (Si), Tin (Sn), Vanadium (V), and Zinc (Zn).</td>
</tr>
</tbody>
</table>

Table 2-1 Principles of Anatomy and Physiology, 11/e  
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How Matter Is Organized: Chemical Elements

(a) Nitrogen deficiency

(b) Iodine deficiency

Figure 2.3 - The effects of essential-elements deficiencies.

a) **The effect of nitrogen deficiency in corn.** In this controlled experiment, the plants on the left are growing in soil that was fertilized with compounds containing nitrogen, while the soil on the right is deficient in nitrogen.

b) **Goiter**, an enlarged thyroid gland, is the result of a deficiency of the trace element iodine. The goiter of this Malaysian woman can probably be reversed by iodine supplements.
How Matter Is Organized: Structure of Atoms

• Atom
  - The *smallest unit of matter* that can enter into a chemical reaction and *retains the properties of an element*. The properties of an element depend on the structure of the element’s atoms.
  - **Hydrogen (H):** the smallest atom; it has a diameter less than 0.1 nanometer \((0.1 \times 10^{-9} \text{ m} = 0.0000000001 \text{ m})\)
  - **Subatomic particles:** Make up an atom.
    - **Protons \((p^+)\):** *Positively* charged particles. The particular type of element is determined by the number of protons.
    - **Neutrons \((n^0)\):** *Uncharged* particles.
    - **Electrons \((e^-)\):** *Negatively* charged particles.
  - **Atomic nucleus:** Dense central *core* where protons and neutrons are localized.

* We symbolize atoms with the same abbreviation used for the element made up of those atoms. For example, \(\text{C}\) stands for both the element carbon and a single carbon atom.
How Matter Is Organized: Structure of Atoms

(a) This model represents the electrons as a cloud of negative charge.

(b) In this model, the electrons are shown as two small blue spheres on a circle around the nucleus.

Figure 2.4 - Simplified models of a helium (He) atom.
* An element’s properties depend on the **structure** of its atoms. *
**Atomic Number, Mass Number, and Atomic Mass**

| **Atomic Number** | Number of *protons* in the nucleus of an atom, unique for each element. 
* All atoms of a particular element have the same number of protons in their nuclei. The atomic number is written as a *subscript* to the left of the symbol for the element. 
✓ Unless otherwise indicated, an atom is neutral in electrical charge, which means that its protons must be balanced by an equal number of electrons. Therefore the atomic number tell us the number of protons and also the number of electrons in an electrically neutral atom. 

| **Example:** | 11\text{Na} 

| **Mass Number** | The sum of *protons and neutrons* in the nucleus of an atom. The mass number is usually written as a *superscript* to the left of the element’s symbol. We can determine the number of neutrons by substracting the atomic number from the mass number. For example, an atom of sodium (Na), has 11 protons, 11 electrons, and 12 neutrons. 

| **Example:** | 23\text{Na} 

| **Atomic Mass or Weight** | Total mass of an atom of a given element (in daltons). The contribution of electrons to mass is negligible. Therefore, almost all of an atom’s mass is concentrated in its nucleus (core). Neutrons and protons each have a mass close to 1 *dalton* (atomic mass unit, or amu), therefore the mass number is an approximation of the total mass of an atom. Example: The atomic mass of sodium (Na) is 23 daltons, although more precisely it is 22.9898 daltons. |
Atoms and Elements: Isotopes

- **Isotopes** are atoms of an element that have different numbers of **neutrons**, and therefore different atomic mass.

- **Example:** Carbon (C):
  - 99% of C in nature is the isotope carbon-12 (\(^{12}\text{C}\)), which has 6 neutrons.
  - The remaining 1% consists of carbon-13 (\(^{13}\text{C}\)) with 7 neutrons, and carbon-14 (\(^{14}\text{C}\)) with 8 neutrons.
  - All isotopes of carbon have 6 protons—otherwise, they would not be carbon.

- Isotopes behave identically in chemical reactions.

Remember that the **atomic number** is the number of **protons** in the nucleus of an atom, and the **mass number** is the sum of **protons** and **neutrons** in the nucleus of an atom.
Atomic Structures of Several Stable Atoms.

Hydrogen (H)
- Atomic number = 1
- Mass number = 1 or 2
- Atomic mass = 1.01

Carbon (C)
- Atomic number = 6
- Mass number = 12 or 13
- Atomic mass = 12.01

Nitrogen (N)
- Atomic number = 7
- Mass number = 14 or 15
- Atomic mass = 14.01

Oxygen (O)
- Atomic number = 8
- Mass number = 16, 17, or 18
- Atomic mass = 16.00

Sodium (Na)
- Atomic number = 11
- Mass number = 23
- Atomic mass = 22.99

Chlorine (Cl)
- Atomic number = 17
- Mass number = 35 or 37
- Atomic mass = 35.45

Potassium (K)
- Atomic number = 19
- Mass number = 39, 40, or 41
- Atomic mass = 39.10

Iodine (I)
- Atomic number = 53
- Mass number = 127
- Atomic mass = 126.90

Atomic number = number of protons in an atom
Mass number = number of protons and neutrons in an atom (boldface indicates most common isotope)
Atomic mass = average mass of all stable atoms of a given element in daltons

Figure 2-2 Principles of Anatomy and Physiology, 11/e
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Atoms and Elements: Isotopes

• **Radioactive isotopes (radioisotopes):** An isotope (atom) that is **unstable**; it **decays** spontaneously, giving off particles and energy (**radioactivity** or **radioactive decay**).
  - Example: The isotope $^{14}\text{C}$ (carbon-14).

• **Weak radioactive isotopes** have many useful applications:
  - Measuring radioactivity in fossils to date those relics of past life.
  - Radioactive tracers help monitor biological processes.
  - Diagnostic and treatment tools in medicine (cancer).

• **Negative effects of radioactivity:**
  - Damage to molecules, cells, and tissues.
  - Environmental: radioactive fallout from nuclear accidents.
Atoms and Elements: Radioactive Isotopes

- **PET**, an acronym for **positron-emission tomography**, detects locations of intense chemical activity in the body. The patient is first injected with a nutrient such as glucose, labeled with a radioactive isotope that emits subatomic particles. These particles collide with electrons made available by chemical reactions in the body. A **PET scanner detects the energy released in these collisions** and maps “hot spots,” the regions of an organ that are most chemically active at the time. The color of the image varies with the amount of the isotope present, with the bright yellow color here identifying a hot spot of cancerous throat tissue.

**FIGURE 2.6** – A PET scan, a medical use for radioactive isotopes.
Structure of Atoms: The Energy Levels of Electrons

• How do atoms interact?
  - When two atoms approach each other during a chemical reaction, their nuclei do not come close enough to interact. *Only electrons are directly involved in the chemical reactions between atoms* (for example, to form compounds).
  - An atom’s electrons vary in the amount of energy they possess (energy is the capacity to cause change, especially to do work).
  - The electrons have potential energy, the energy that matter possesses because of its location or structure.
  - The negatively charged electrons are attracted to the positively charged protons in the atomic nucleus. It takes work to move an electron farther away from the nucleus.
  - In an atom, electrons occupy specific energy levels, each of which can be represented by an electron shell of that atom.
Structure and Interaction of Atoms: The Energy Levels of Electrons

- **Electron shell**: A region within which electrons orbit that corresponds to a fixed *energy level* at a given distance from the nucleus of an atom (somewhat like planets orbiting the sun). In each level, the electrons occupy *orbitals* (volumes of space around the atomic nucleus).

- The electrons usually fill the shell closest to the nucleus first, and then begin to occupy the next shell. *The first electron shell can hold no more than 2 electrons. The second shell holds a maximum of 8 electrons.*

- *The chemical reactivity* of an atom depends on the number of electrons (*valence electrons*) in its outermost electron shell (*valence shell*); an atom is most stable, and therefore least reactive, when its outermost shell is either completely full or empty.

**Figure 2.1 Principles of Anatomy and Physiology, 11/e © 2006 John Wiley & Sons**

***An atom that has an incomplete valence shell is *reactive*. In other words, the reactivity of an atom arises from the presence of unpaired electrons in the valence shell. *Atoms interact in a way that completes their valence shells.*
Structure of Atoms: The Energy Levels of Electrons

a) A ball bouncing down a flight of stairs provides an analogy for energy levels of electrons, because the ball can come to rest only on each step, not between steps.

b) An electron can move from one shell to another only if the energy it gains or loses is exactly equal to the difference in energy between the energy levels of the two shells. Arrows indicate some of the stepwise changes in potential energy that are possible.
The Periodic Table of Elements

• The *Periodic Table of the Elements* is a tabular arrangement of the elements according to their *atomic numbers* so that elements with similar properties are in the same vertical column (*group*).

  ➢ The elements are arranged in rows, or *periods*, corresponding to the number of electron shells in their atoms.

  ➢ The *left-to-right sequence* of elements in each row corresponds to the sequential addition of electrons and protons. (See the complete Periodic Table of the Elements).
Figure 2-1. The Periodic Table

- **Chemical symbol**
- **Atomic number**
- **Chemical name**

Number of e⁻ in each energy level:
- Na: 11
- Mg: 12
- C: 6
- H: 1

Atomic Masses (amu):
- H: 1.01 amu
- Mg: 24.31 amu
- C: 12.01 amu
- He: 4.00 amu

Elements:
- Hydrogen (H)
- Magnesium (Mg)
- Carbon (C)
- Sodium (Na)
- Neon (Ne)
- Potassium (K)
- Chlorine (Cl)

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In a standard periodic table, information for each element is presented as shown for helium in the inset. In this diagram, the electrons are represented as blue dots and electron shells as concentric circles. The elements are arranged in rows, each representing the filling of an electron shell. As electrons are added, they occupy the lowest available shell.
# BIOLOGY I. Chapter 2 – The Chemical Context of Life

## PERIODIC TABLE OF THE ELEMENTS

<table>
<thead>
<tr>
<th>Group</th>
<th>Period</th>
<th>Element</th>
<th>Atomic Number</th>
<th>Symbol</th>
<th>Molar Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>IIA</td>
<td>1</td>
<td>Hydrogen</td>
<td>1</td>
<td>H</td>
<td>1.0079</td>
</tr>
<tr>
<td>IIA</td>
<td>2</td>
<td>Lithium</td>
<td>3</td>
<td>Li</td>
<td>6.9412</td>
</tr>
<tr>
<td>IIA</td>
<td>3</td>
<td>Sodium</td>
<td>11</td>
<td>Na</td>
<td>22.990</td>
</tr>
<tr>
<td>IA</td>
<td>1</td>
<td>Potassium</td>
<td>19</td>
<td>K</td>
<td>39.098</td>
</tr>
<tr>
<td>IA</td>
<td>2</td>
<td>Rubidium</td>
<td>37</td>
<td>Rb</td>
<td>85.465</td>
</tr>
<tr>
<td>IA</td>
<td>3</td>
<td>Caesium</td>
<td>55</td>
<td>Cs</td>
<td>132.91</td>
</tr>
<tr>
<td>IA</td>
<td>4</td>
<td>Francium</td>
<td>87</td>
<td>Fr</td>
<td>223.01</td>
</tr>
<tr>
<td>IIA</td>
<td>1</td>
<td>Helium</td>
<td>2</td>
<td>He</td>
<td>4.0026</td>
</tr>
<tr>
<td>IIA</td>
<td>2</td>
<td>Oxygen</td>
<td>16</td>
<td>O</td>
<td>15.999</td>
</tr>
<tr>
<td>IIA</td>
<td>3</td>
<td>Neon</td>
<td>10</td>
<td>Ne</td>
<td>20.182</td>
</tr>
</tbody>
</table>

---

**Note:**

2. The values enclosed in brackets indicate the mass number of the longest-lived isotope of the element.
3. However, three such elements (Th, Pm, and U) do have a characteristic terrestrial isotopic composition, and for these their atomic weights are tabulated.

**Editor:** Adriana Vothan (adivov@netinx.com)
How Matter Is Organized: Ions, Molecules, and Compounds

- **Ion**
  - Any atom or molecule with an *electric charge*, either *positive or negative*, due to unequal numbers of protons and electrons (it is usually formed when a substance, such as salt, dissolves and dissociates).

- **Molecule**
  - The chemical combination of two or more atoms held together by a chemical bond (a force of attraction).
    - *Molecular formula*: Indicates the elements and the number of atoms of each element that make up a molecule. Example: Oxygen is $O_2$.

- **Compound**
  - A substance consisting of two or more *different elements* combined in a fixed ratio; it can be broken down into two or more other substances by chemical means.
    - Examples: carbon dioxide ($CO_2$), and water ($H_2O$).
<table>
<thead>
<tr>
<th>Common name</th>
<th>Water</th>
<th>Familiar term.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chemical name</td>
<td>Hydrogen oxide</td>
<td>Systematically describes elemental composition.</td>
</tr>
<tr>
<td>Chemical formula</td>
<td>H$_2$O</td>
<td>Indicates unvarying proportions of elements. Subscripts show number of atoms of an element per molecule. The absence of a subscript means one atom.</td>
</tr>
<tr>
<td>Structural formula</td>
<td>H–O–H</td>
<td>Represents each covalent bond as a single line between atoms. The bond angles also may be represented.</td>
</tr>
<tr>
<td>Structural model</td>
<td><img src="image" alt="Structural model" /></td>
<td>Shows the positions and relative sizes of atoms.</td>
</tr>
<tr>
<td>Shell model</td>
<td><img src="image" alt="Shell model" /></td>
<td>Shows how pairs of electrons are shared in covalent bonds.</td>
</tr>
</tbody>
</table>
Compounds, such as sodium chloride (NaCl), contain atoms of two or more different elements and can be broken down into two or more other substances by chemical means.
How Matter Is Organized: Free Radicals

- Free radical
  - Electrically charged atom or group of atoms with an unpaired electron in its outermost shell.
  - Free radicals are produced by some normal metabolic reactions of the body as well as by radiation (ultraviolet and X-rays) and by harmful chemicals. They are unstable, highly reactive, and destructive to nearby molecules.
    - Example: superoxide (O$_2^-$).
  - Antioxidants are chemicals that neutralize free radicals (example: vitamins C and E).
Chemical Bonds

- A chemical bond is the force of attraction that holds together the atoms of a molecule or a compound.
- The formation and function of molecules depend on chemical bonding between atoms.
  - The valence of an atom is its bonding capacity.
  - The likelihood that an atom will form a chemical bond with another atom depends on the number of electrons in its outermost shell: the valence shell.
  - The strongest kinds of chemical bonds are covalent bonds. Other types are ionic bonds and hydrogen bonds.
Chemical Bonds

• Covalent Bonds
  - **Sharing** of two or more electrons rather than gaining or losing whole electrons. Covalent bonds are **stronger** than other types of chemical bonds.
  - The **larger the number of electron pairs** shared between two atoms, the **stronger** the covalent bond.
    - **Single covalent bond**: sharing a pair of electrons.
    - **Double covalent bond**: sharing two pairs of electrons.
  - Examples: Covalent bonds in a molecule of hydrogen ($\text{H}_2$), and in a molecule of water ($\text{H}_2\text{O}$).
**DIAGRAMS OF ATOMIC AND MOLECULAR STRUCTURE**

(a) **Hydrogen atoms**

\[ \text{H} + \text{H} \rightarrow \text{H} - \text{H} \]

**Hydrogen molecule**

\[ \text{H}_2 \]

(b) **Oxygen atoms**

\[ \text{O} + \text{O} \rightarrow \text{O} = \text{O} \]

**Oxygen molecule**

\[ \text{O}_2 \]

(c) **Nitrogen atoms**

\[ \text{N} + \text{N} \rightarrow \text{N} \equiv \text{N} \]

**Nitrogen molecule**

\[ \text{N}_2 \]

(d) **Carbon atom**

\[ \text{C} + 4\text{H} \rightarrow \text{CH}_4 \]

**Methane molecule**

\[ \text{CH}_4 \]

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*Figure 2-5 Principles of Anatomy and Physiology, 11/e © 2006 John Wiley & Sons*
Chemical Bonds

• Covalent Bonds

- The attraction of a particular kind of atom for the electrons of a covalent bond is called its **electronegativity**.
  - *The greater the number of protons, the greater the electronegativity*. The more electronegative an atom, the more strongly it pulls shared electrons toward itself.

- **Nonpolar covalent bond**: electrons are *shared equally* between two atoms of similar electronegativity.

- **Polar covalent bond**: electrons are *not shared equally* between atoms that differ in electronegativity; the shared electrons are pulled closer and stronger to the more electronegative atom, making it slightly negative and the other atom slightly positive.
Chemical Bonds: Covalent Bonds

Water is a polar molecule. The electrons (red) are shared unequally. Because the oxygen nucleus attracts the shared electrons more strongly, the oxygen end of a water molecule has a partial negative charge; written $\delta^-$, and the hydrogen ends have partial positive charges, written $\delta^+$. 

Oxygen atom  Hydrogen atoms  Water molecule

Figure 2-6 Principles of Anatomy and Physiology, 11/e
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Chemical Bonds

• Ionic Bonds
  - Force of attraction that holds ions (charged atoms or molecules) having opposite charges together. Two atoms are so unequal in their attraction for valence electrons that the more electronegative atom strips an electron completely away from its partner.
  - Cation: A positively (+) charged ion.
    - Example: A sodium ion (Na\(^+\)).
  - Anion: A negatively (-) charged ion.
    - Example: A chloride ion (Cl\(^-\)).
  - Compounds formed by ionic bonds are called ionic compounds or salts.
  - Ionic bonds are weaker than covalent bonds.
Chemical Bonds: Ionic Bonds

- **An ionic bond is the force of attraction that holds together oppositely charged ions.** The more electronegative atom strips an electron completely away from its partner.

(a) A sodium atom (Na) can have a complete octet of electrons in its outermost shell by **losing** one electron (Because sodium no longer has an electron in the third shell, the second shell is now the valence shell).

(b) A chlorine atom (Cl) can have a complete octet in its valence shell by **gaining** one electron.

(c) An ionic bond may form between oppositely charged ions.

(d) In a crystal of NaCl (table salt), each Na\(^+\) is surround by six Cl\(^-\). In (a), (b) and (c), the electron that is lost or accepted is colored red.
After the electron transfer between the two atoms, sodium has 11 protons but only 10 electrons, for an electrical charge of 1+. Chlorine has 17 protons and 18 electrons, giving it a net electrical charge of 1-.

During the formation of sodium chloride, an electron is transferred from the sodium atom to the chlorine atom. At the completion of the reaction, each atom has eight electrons in the outer shell, but each also carries a charge as shown.

In a sodium chloride crystal, ionic bonding between Na+ and Cl- causes the atoms to assume a three-dimensional lattice in which each sodium ion is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions. The result is crystals of salt as in table salt.

**Figure 2.7. Formation of sodium chloride (table salt).**
Chemical Bonds: Ionic Bonds (Animation)

- **Sodium (Na)**, with only one electron in its third shell, tends to be an *electron donor*. Once it gives up this electron, the second shell, with eight electrons, becomes its outer shell (valence shell). **Chlorine (Cl)**, on the other hand, tends to be an *electron acceptor* (it is more *electronegative*). Its outer shell has seven electrons, so if it acquires only one more electron it has a completed outer shell. When a sodium atom and a chlorine atom come together, *an electron is transferred from the sodium atom to the chlorine atom*. Now both atoms have eight electrons in their outer shells.
# TABLE 2.2 Common Ions and Ionic Compounds in the Body

<table>
<thead>
<tr>
<th>Cations</th>
<th>Symbol</th>
<th>Anions</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Name</td>
<td></td>
<td>Name</td>
<td></td>
</tr>
<tr>
<td>Hydrogen ion</td>
<td>H⁺</td>
<td>Fluoride ion</td>
<td>F⁻</td>
</tr>
<tr>
<td>Sodium ion</td>
<td>Na⁺</td>
<td>Chloride ion</td>
<td>Cl⁻</td>
</tr>
<tr>
<td>Potassium ion</td>
<td>K⁺</td>
<td>Iodide ion</td>
<td>I⁻</td>
</tr>
<tr>
<td>Ammonium ion</td>
<td>NH₄⁺</td>
<td>Hydroxide ion</td>
<td>OH⁻</td>
</tr>
<tr>
<td>Hydronium ion</td>
<td>H₃O⁺</td>
<td>Nitrate ion</td>
<td>NO₃⁻</td>
</tr>
<tr>
<td>Magnesium ion</td>
<td>Mg²⁺</td>
<td>Bicarbonate ion</td>
<td>HCO₃⁻</td>
</tr>
<tr>
<td>Calcium ion</td>
<td>Ca²⁺</td>
<td>Oxide ion</td>
<td>O²⁻</td>
</tr>
<tr>
<td>Iron (II) ion</td>
<td>Fe²⁺</td>
<td>Sulfate ion</td>
<td>SO₄²⁻</td>
</tr>
<tr>
<td>Iron (III) ion</td>
<td>Fe³⁺</td>
<td>Phosphate ion</td>
<td>PO₄³⁻</td>
</tr>
</tbody>
</table>

Table 2-2 Principles of Anatomy and Physiology, 11/e  
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Weak Chemical Bonds

- Hydrogen Bonds
  - A hydrogen atom with a partial positive charge ($\delta^+$) attracts the partial negative charge ($\delta^-$) of neighboring electronegative atoms, most often oxygen or nitrogen in living cells.
  - Among water molecules ($H_2O$).
  - Hydrogen bonds are weak, so they form and break easily; but they do establish important links between molecules or between different parts of biological molecules, such as proteins or nucleic acids.
    - Example: Hydrogen bonds are found between paired nitrogenous bases of the DNA molecule (a nucleic acid).
Chemical Bonds: **Hydrogen Bonds**

- **Hydrogen bonds** among water molecules occur because hydrogen atoms in one water molecule are attracted to the partial negative charge of the oxygen atom in another water molecule.
Chemical Reactions

- A chemical reaction is the combination or separation of atoms in which chemical bonds are formed or broken and new products with different properties are produced.
  - Chemical reactions are the foundation of all life processes.
  - Reactants: Starting substances.
  - Products: Ending substances.

- Metabolism is all of the reactions in the cells and tissues of the body.
Chemical Reactions

\[ 2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O} \]

**Reactants**  
2 H\(_2\) \hspace{2cm} \text{O}_2 

**Products**
2 H\(_2\)O

**FIGURE 2.8 - A Chemical Reaction.**

- The chemical reaction between two hydrogen molecules (H\(_2\)) and one oxygen molecule (O\(_2\)) to form two molecules of water (H\(_2\)O). Note that the reaction occurs by breaking old bonds and making new bonds.
- The number of atoms of each element is the same before and after a chemical reaction.
Forms of Energy and Chemical Reactions

• Energy
  ➢ Capacity to cause change, especially to do work.

• Potential energy
  ➢ Energy stored by matter due to its position.

• Kinetic energy
  ➢ Energy associated with matter in motion.

• Chemical energy
  ➢ Form of potential energy that is stored in the bonds of compounds and molecules.
Energy Transfer in Chemical Reactions

- **Exergonic reactions**
  - **Release** more energy than they absorb.

- **Endergonic reactions**
  - **Absorb** more energy than they release.
Types of Chemical Reactions

- **Synthesis Reactions—Anabolism**
  - Combination of two or more atoms, ions, or molecules to form new and larger molecules.
  
  \[ A + B \rightarrow AB \]

- **Anabolism**: All the synthesis reactions that occur in the body of an organism.
Types of Chemical Reactions

• Decomposition Reactions—**Catabolism**
  - Splitting up of large molecules into smaller molecules, ions, or atoms.

    \[ AB \rightarrow A + B \]

  - **Catabolism**: All the decomposition reactions that occur in the body of an organism.
Types of Chemical Reactions

- **Exchange Reactions**
  - They consist of both synthesis and decomposition reactions.
  
  \[
  AB + CD \rightarrow AD + BC
  \]

- **Reversible Reactions**
  - The products can revert to the original reactants.
    
    \[
    AB \leftrightarrow A + B
    \]
Introduction to Inorganic and Organic Compounds

- **Inorganic Compounds (or Molecules)**
  - Compounds **without carbon**; they usually are small and **simple** (and often contain ionic bonds).
  - Examples: water and many acids, bases, and salts.
  - **Carbon dioxide (CO₂)** is an exception; although it contains carbon, it is considered inorganic because it is very simple and lacks hydrogen.

- **Organic Compounds (or Molecules)**
  - Always **contain carbon**, usually contain **hydrogen**, and are typically structurally **complex** (and usually contain covalent bonds).
  - Examples: carbohydrates, lipids, proteins, nucleic acids.
 Ion = An atom or molecule that carries an electrical charge (positive or negative).
 Ionic bond = Force of attraction that holds ions having opposite charges together.
 Covalent bond = Sharing of two or more electrons rather than gaining or losing whole electrons.

**TABLE 3.1**

<table>
<thead>
<tr>
<th>Inorganic Molecules</th>
<th>Organic Molecules</th>
</tr>
</thead>
<tbody>
<tr>
<td>Usually contain positive and negative ions</td>
<td>Always contain carbon and hydrogen</td>
</tr>
<tr>
<td>Usually ionic bonding</td>
<td>Always covalent bonding</td>
</tr>
<tr>
<td>Always contain a small number of atoms</td>
<td>Often quite large, with many atoms</td>
</tr>
<tr>
<td>Often associated with nonliving matter</td>
<td>Usually associated with living organisms</td>
</tr>
</tbody>
</table>
“Perhaps one of you gentlemen would mind telling me just what it is outside the window that you find so attractive...?”

Another casualty in the War of the Atoms.
References


