Chapter 2

2.1 Chemical Elements

1. **Matter** is defined as anything that takes up space and has mass.
2. Matter exists in three states: **solid, liquid, and gas**.

A. Elements

1. All matter (both living and non-living) is composed of 92 naturally-occurring **elements**.
2. Elements, by definition, cannot be broken down to simpler substances with different chemical or physical properties.
3. Six elements (carbon, hydrogen, nitrogen, oxygen, phosphorus, and sulfur—acronym CHNOPS) make up 95% of the body weight of organisms.

B. Atoms

1. Elements consist of tiny particles called **atoms.**
2. An atom is the smallest unit of an element that displays the properties of the element.
3. One or two letters (e.g., H, Na) create the **atomic symbol** of the element
4. Atoms contain specific numbers of **protons, neutrons,** and **electrons.**
5. Protons are positively charged particles; neutrons have no charge; electrons are negatively charged particles located in orbitals outside the nucleus.
6. Protons and neutrons are in the nucleus of an atom; electrons move around the nucleus.

C. Atomic Mass and Atomic Number

1. The **atomic mass** of an atom depends on the presence of certain subatomic particles.
2. Atomic mass is the sum of protons and neutrons
3. Protons and neutrons have one atomic mass unit (amu of weight); electrons have zero.
4. All atoms of an element have the same number of protons, called the **atomic number** of the element.

D. The Periodic Table

1. The periodic table shows how various characteristics of atoms of elements recur.
2. **Groups** are the vertical columns in the table, **periods** are the horizontal rows; atomic mass increases as you move down a group or across a period.
3. The atomic number is above the atomic symbol and the atomic mass is below the atomic symbol.

E. Isotopes

1. **Isotopes** are atoms of the same element that differ in the number of neutrons (and therefore have different atomic masses). For example, carbon-12 has 6 protons and 6 neutrons, carbon-14 has 6 protons and 8 neutrons.

2. A carbon atom with 8 rather than 6 neutrons is unstable; it releases energy and subatomic particles and is thus a **radioactive isotope**.
3. Because the chemical behavior of a radioactive isotope is the same as a stable isotope of a particular element, low levels of the radioactive isotope (e.g., radioactive iodine or glucose) allow researchers to trace the location and activity of the element in living tissues; these isotopes are called **tracers**.
4. High levels of radiation can destroy cells and cause cancer; careful use of radiation can sterilize products and kill cancer cells.

F. Electrons and Energy

1. Electrons occupy orbitals within various **energy levels** (or **electron shells**) near or distant from the nucleus of the atom. The farther the orbital from the nucleus, the higher the energy level.
2. An orbital is a volume of space where an electron is most likely to be found; an orbital can contain no more than 2 electrons.
3. When atoms absorb energy during photosynthesis, electrons are boosted to higher energy levels. When the electrons return to their original energy level, the released energy is converted into chemical energy. This chemical energy supports all life on Earth.
4. The innermost shell of an atom is complete with 2 electrons; all other shells are complete with 8 electrons. This is called the **octet rule**.
5. Atoms will give up, accept, or share electrons in order to have 8 electrons in an electron shell.

2.2 Compounds and Molecules

1. When atoms of two or more different elements bond together, they form a **compound** (e.g., H_2O).
2. A **molecule** is the smallest part of a compound that has the properties of the compound.
3. A **formula** tells you the number of each kind of atom in a molecule (ex. Glucose, $C_6H_{12}O_6$)
4. Electrons possess energy, and bonds that exist between atoms in molecules therefore contain energy.

A. Ionic Bonding

1. An **ionic bond** forms when electrons are transferred from one atom to another atom.
2. By losing or gaining electrons, atoms fill outer shells, and are more stable (the octet rule).
3. Example: sodium loses an electron and therefore has a positive charge; chlorine gains an electron to give it a negative charge. Such charged particles are called **ions.**
4. Attraction of oppositely charged ions holds the two atoms together in an **ionic bond**.
5. A **salt** (e.g., NaCl) is an example of an ionically-bonded compound.

B. Covalent Bonding

1. **Covalent bonds** result when two atoms share electrons so each atom has an octet of electrons in the outer shell (or, in the case of hydrogen, 2 electrons).
2. Hydrogen can give up an electron to become a hydrogen ion (H^+) or share an electron with another atom to complete its shell with 2 electrons.
3. The **structural formula** of a compound indicates a shared pair of electrons by a line between the two atoms; e.g., single covalent bond (H–H), double covalent bond (O=O), and triple covalent bond ($N \equiv N$). Each line between the atoms represents a pair of electrons.

C. Nonpolar and Polar Covalent Bonds

1. In **nonpolar covalent bonds**, sharing of electrons is equal, i.e., the electrons are not attracted to either atom to a greater degree.
2. With **polar covalent bonds**, the sharing of electrons is unequal.
   a. In a water molecule (H_2O), sharing of electrons by oxygen and hydrogen is not equal; the oxygen atom with more protons attracts the electrons closer to it, and thus dominates the H_2O association.
   b. Attraction of an atom for electrons in a covalent bond is called the **electronegativity** of the atom**;** an oxygen atom is more electronegative than a hydrogen atom.
   c. Oxygen in a water molecule, more attracted to the electron pair, assumes a partial negative charge.

2.3. Chemistry of Water

1. The shape of water and of all organic molecules is necessary to the structural and functional roles they play in living things.
2. A **hydrogen bond** is the attraction of a slightly positive hydrogen to a slightly negative atom in the vicinity.

A. Hydrogen Bonding

1. A **hydrogen bond** is a weak attractive force between the slightly positive charge of the hydrogen atom of one molecule and slightly negative charge of another atom (e.g., oxygen, nitrogen) in another or the same molecule.
2. Many hydrogen bonds taken together are relatively strong.
3. Hydrogen bonds between and within complex biological molecules (e.g., DNA, proteins) help maintain their proper structure and function.

B. Properties of Water

1. Water has a **high heat capacity**
   a. The temperature of liquid water rises and falls more slowly than that of most other liquids.
   b. A **calorie** is the amount of heat energy required to raise the temperature of one gram of water 1° C.
   c. Because the hydrogen bonds between water molecules hold more heat, water's temperature falls more slowly than other liquids; this protects organisms from rapid temperature changes and helps them maintain homeostatic temperature.

2. Water has a high **heat of evaporation**.
   a. When water boils, it evaporates, or vaporizes into the environment.
   b. Hydrogen bonds between water molecules require a relatively large amount of heat to break.
   c. This property moderates Earth's surface temperature; permits living systems to exist.
   d. When animals sweat, evaporation of the sweat removes body heat, thus cooling the animal.
3. Water is a **solvent.**
   a. Water dissolves a great number of substances (e.g., salts, large polar molecules).
   b. A **solution** contains dissolved substances called **solutes.**
   c. Ionized or polar molecules attracted to water are **hydrophilic** ("water loving")**.**
   d. Nonionized and nonpolar molecules that cannot attract water are **hydrophobic** ("water fearing").
4. Water molecules are **cohesive** and **adhesive**.
   a. **Cohesion** allows water to flow freely without molecules separating.
   b. **Adhesion** is ability to adhere to polar surfaces; water molecules have positive and negative poles.
   c. Water rises up a tree from roots to leaves through small tubes.
      1) Adhesion of water to walls of vessels prevents water column from breaking apart.
      2) Cohesion allows evaporation from leaves to pull water column from roots.
   d. Water has a **high surface tension** and is relatively difficult to break through at its surface.
      1) This property permits a rock to be skipped across a pond surface, and supports insects walking on surface.
5. Unlike most substances, **frozen water (ice) is less dense than liquid water**.
   a. Below 4º C, hydrogen bonding becomes more rigid but more open, causing expansion.
   b. Because ice is less dense, it floats; therefore, bodies of water freeze from the top down.
   a. If ice was heavier than water, ice would sink and bodies of water would freeze solid.
   b. This property allows ice to act as an insulator on bodies of water, thereby protecting aquatic organisms during the winter.

2.4. Acids and Bases

1. When water ionizes or dissociates, it releases a small (1×10^7 moles/liter) but equal number of hydrogen (H^+) ions and hydroxide (OH^-) ions; $H - O - H \rightarrow H^+ + OH^-$.

2. **Acid** molecules dissociate in water, releasing hydrogen (H^+) ions: $HCl \rightarrow H^+ + Cl^-$.
3. **Bases** are molecules that take up hydrogen ions or release hydroxide ions. $NaOH \rightarrow Na^+ + OH^-$.
4. The **pH scale** indicates acidity and basicity (alkalinity) of a solution.
   a. **pH** is the measurement of free hydrogen ions, expressed as a negative logarithm of the H^+ concentration ($-\log [H^+]$).
   b. **pH** values range from 0 (1×10^0 moles/liter; most acidic) to 14 (1×10^{-14} moles/liter; most basic).
      1) One mole of water has 1×10^{-7} moles/liter of hydrogen ions; therefore, has neutral pH of 7.
      2) An acid is a substance with pH less than 7; a base is a substance with pH greater than 7.
      3) Because it is a logarithmic scale, each lower unit has 10 times the amount of hydrogen ions as next higher pH unit; as move up pH scale, each unit has 10 times the basicity of previous unit.
5. **Buffers** keep pH steady and within normal limits in living organisms..
   a. Buffers stabilize pH of a solution by taking up excess hydrogen (H^+) or hydroxide (OH^-) ions.
   b. Carbonic acid helps keep blood pH within normal limits: $H_2CO_3 \rightarrow H^+ + HCO_3$-.

A. The Harm Done by Acid Deposition (*Ecology Focus box*)
   1. **Acid deposition** is rain or snow with pH < 5.0, and dry acidic particles that settle on the Earth from the atmosphere.
   2. The burning of fossil fuels increases the amount of acid deposition that falls from the atmosphere to the Earth.
   3. Impact of Lakes
      a. Aluminum is leached from the soil converts mercury deposits to methyl mercury, which is toxic and accumulates in fish.
   4. Impact on Forests
      a. Acid rain damages plant leaves so they cannot conduct photosynthesis.
      b. Acid rain stresses plants and they are then more susceptible to disease and pests.
      c. When toxins are leached into the soil, the toxins can kill vital soil fungi that help roots absorb nutrients.
   5. Impact on Humans and Structures
      a. Inhaling dry acidic particles can increase the chance of respiratory illnesses such as asthma.