# *CHAPTER* 2 BASIC CHEMISTRY

## Chapter Outline

**2.1 Chemical Elements**

A. Matter
   1. Matter takes up space and has mass.
   2. All living and nonliving matter is composed of 92 naturally-occurring basic elements.
   3. Elements cannot be broken down to substances with different chemical or physical properties.
   4. Six elements (C, H, N, O, P, S) make up 98% of living things.

B. Atomic Structure
   1. Chemical and physical properties of atoms (e.g., mass) depend on the subatomic particles.
      a. Different atoms contain specific numbers of **protons, neutrons,** and **electrons.**
      b. Protons and neutrons are in nucleus of atoms; electrons move around nucleus.
      c. **Protons** are positively charged particles; neutrons have no charge; both have about 1 atomic mass unit of weight.
      d. **Electrons** are negatively charged particles.
   2. The **atomic mass** of an atom is about equal to the sum of its protons and neutrons.
   3. All atoms of an element have the same number of protons, the atom's **atomic number**.

C. The Periodic Table
   1. The periodic table shows how various characteristics of atoms recur.
   2. The table is arranged in order of atomic number, with periods in horizontal rows and groups in vertical columns.

D. Isotopes
   1. **Isotopes** are atoms with the same number of protons but differ in number of neutrons; e.g., a carbon atom has six protons but may have more or less than usual six neutrons.
   2. A carbon with eight rather than six neutrons is unstable; it releases rays and subatomic particles and is a radioactive isotope.
   3. Low levels of radiation such as radioactive iodine or glucose allow researchers to trace the location and activity of the atom in living tissues; therefore these isotopes are called "tracers."
   4. High levels of radiation can cause cancerous tissues and destroy cells; careful use of radiation in turn can sterilize products and kill cancer cells.

E. Electrons and Energy
   1. Electrons occupy an orbital at some level near or distant from the nucleus of the atom.
   2. An orbital is a volume off space where an electron is most likely to be found; an orbital contains no more than two electrons.
   3. The more distant the orbital, the more energy it takes to stay in the orbital.
   4. When atoms absorb energy during photosynthesis, electrons are boosted to higher energy levels.
   5. The innermost shell of an atom is complete with two electrons; all other shells are complete with eight electrons.

**2.2 Elements and Compounds**

A. Compounds
   1. When two or more different elements react or 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. Electrons possess energy and bonds that exist between atoms in molecules contain energy.

B. **Ionic Bonding**
   1. Ionic bonds form when electrons are transferred from one atom to another.
   2. Losing or gaining electrons, atoms participating in ionic reactions fill outer shells, and are more stable.
   3. Example: sodium with one less electron has positive charge; chlorine has extra electron that has negative charge. Such charged particles are called **ions.**
   4. Attraction of oppositely charged ions holds the two atoms together in an **ionic bond**.

C. **Covalent Bonding**
   1. Covalent bonds result when two atoms share electrons so each atom has octet of electrons in the outer shell.
   2. Hydrogen can give up an electron to become a hydrogen ion ($H^+$) or share an electron with another atom to complete its outer shell of two electrons.
   3. **Structural formulas** represent shared atom as a line between two atoms; e.g., single covalent bond (H–H), double covalent bond (O=O), and triple covalent bond ($N \equiv N$).
   4. Three dimensional shape of molecules is not represented by structural formulas but shape is critical in understanding the biological action of molecules: action of insulin, HIV receptors, etc.

D. Nonpolar and Polar Covalent Bonds
   1. In **nonpolar covalent bonds**, sharing of electrons is equal.
   2. With **polar covalent bonds**, the sharing of electrons is unequal.
      a. In water molecule ($H_2O$), sharing of electrons by oxygen and hydrogen is not equal; the oxygen atom with more protons dominates the $H_2O$ association.
      b. Attraction of an atom for electrons in a covalent bond is called **electronegativity;** oxygen atom is more electronegative than hydrogen atom.
      c. Oxygen in water molecule, more attracted to electron pair, assumes small negative charge.

E. Hydrogen Bonding
   1. A **hydrogen bond** is weak attractive force between slightly positive hydrogen atom of one molecule and slightly negative atom in another or the same molecule.
   2. Many hydrogen bonds taken together are relatively strong.
   3. Hydrogen bonds between complex molecules of cells help maintain structure and function.

**2.3. Chemistry of Water**

A. First Cells Evolved in Water
   1. All living things are 70–90% water.
   2. Because water is a polar molecule, water molecules are hydrogen bonded to each other.
   3. With hydrogen bonding, water is liquid between 0º C and 100 C which is critical for life.

B. Properties of Water
   1. The temperature of liquid water rises and falls more slowly than that of most other liquids..
      a. **Calorie** is amount of heat energy required to raise temperature of one gram of water $1^{o}$ C.
      b. Because water holds more heat, its temperature falls more slowly than other liquids; this protects organisms from rapid temperature changes and helps them maintain normal temperatures.
   2. Water has a high **heat of vaporization**.
      a. Hydrogen bonds between water molecules require a large amount of heat to break.
      b. This property moderates earth's surface temperature; permits living systems to exist here.
      c. When animals sweat, evaporation of the sweat takes away body heat, thus cooling the animal.
   3. Water is universal solvent, facilitates chemical reactions both outside of and within living systems..
      a. Water is a **universal solvent** because it dissolves a great number of solutes.
      b. Ionized or polar molecules attracted to water are **hydrophilic.**
      c. Nonionized and nonpolar molecules that cannot attract water are **hydrophobic**.
   4. Water molecules are cohesive and adhesive..

a. **Cohesion** allows water to flow freely without molecules separating, due to hydrogen bonding.

b. **Adhesion** is ability to adhere to polar surfaces; water molecules have positive, negative poles.

c. Water rises up 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.

5. Water has a high surface tension measured by how difficult it is to break the surface of a liquid..

a. As with cohesion, hydrogen bonding causes water to have high surface tension.

b. Permits a rock to be skipped across pond surface; supports insect walking on water surface.

6. Unlike most substances, frozen water is less dense than liquid water. .

a. Below $4^o$ 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.

c. If ice was heavier than water, ice would sink and ponds would freeze solid.

C. Acids and Bases

1. Covalently bonded water molecules ionize; the atoms dissociate into ions.
2. When water ionizes or dissociates, it releases a small ($10^7$ moles/liter) but equal number of $H^+$ and OH ions; thus, its *pH* is *neutral*.
3. Water dissociates into hydrogen and hydroxide ions: H – O –H $\rightarrow$ $H^+ + OH^-$.
4. **Acid** molecules dissociate in water, releasing hydrogen ions ($H^+$) ions: $HCl \rightarrow H^+ + Cl^-$.
5. **Bases** are molecules that take up hydrogen ions or release hydroxide ions. $NaOH \rightarrow Na^+ + OH^-$.
6. The **pH scale** indicates acidity and basicity (alkalinity) of a solution. (Fig. 2.13)

a. Measure of free hydrogen ions as a negative logarithm of the $H^+$ concentration (-log $[H^+]$).

b. **pH** values range from 0 ($10^0$ moles/liter; most acidic) to 14 ($10^{14}$ moles/liter; most basic).

**1)** One mole of water has $10^7$ moles/liter of hydrogen ions; therefore, has neutral pH of 7.

**2)** Acid is a substance with pH less than 7; base is a substance with pH greater than 7.

**3)** As 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.

7. **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: $H_2CO_3 \longrightarrow H^+ + HCO_3$-.