Lecture 2: Chemical Basis of Life

I. Background
Chemistry is fundamental to understanding life. All matter is comprised of mass.

Chemical Elements and Compounds
1. Chemical elements, in pure and compound form, comprise matter

Terms:
Matter—anything having mass and occupying space
Mass—a measure of the amount of matter that an object contains

Mass ≠ Weight
Weight is the measure of how strongly an object is pulled by earth’s gravity and consequently varies as a function of distance from the earth’s center. Mass does not vary with its position.

2. Biologically Important Elements
   a. 92 naturally occurring elements
   b. 25 are essential for life
      i. C, O, H and N—96% of all living matter
      ii. Ca, P, K, S, Na, Cl, Mg, and trace elements—remaining 4%
      iii. Trace elements—although in very low quantities, are absolutely essential;
           B, Cr, Co, Cu, F, I, Fe, Mn, Mo, Se, Si, Sn, V and Zn

3. Compounds—two or more compounds combined in a fixed ratio
   a. Compounds have unique emergent properties (NaCl has properties that differ from Na and Cl)

Atoms and Molecules
Terms:
Atom—smallest possible unit of matter that retains the physical and chemical properties of its element; all atoms of the same element share the same chemical properties

1. Atomic structure determines the behavior of an element
   a. Atoms are comprised of subatomic properties (many different types; three that are relatively stable)
      i. Neutrons
      ii. Protons
      iii. Electrons
   
   b. Subatomic Properties

<table>
<thead>
<tr>
<th>Charge</th>
<th>Neutron</th>
<th>Proton</th>
<th>Electron</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>No Charge</td>
<td>+1</td>
<td>-1</td>
</tr>
<tr>
<td>Location</td>
<td>Nucleus</td>
<td>Nucleus</td>
<td>Orbits Nucleus</td>
</tr>
<tr>
<td>Mass</td>
<td>1 dalton</td>
<td>1 dalton</td>
<td>1/2000 dalton</td>
</tr>
</tbody>
</table>

   i. Nucleus—dense core of the atom
ii. Nucleus is positively charged—neutrons and protons (positively charged) only
iii. Electrons are held in orbits by the electrostatic attraction to nuclear protons
iv. Dalton—unit of mass equal to $1.67 \times 10^{-24}$ g

**Atomic Number**

Atomic number—equal to the number of protons in a particular element

a. All atoms of a particular element have the same atomic number
b. Nomenclature: $^{11}\text{Na}$ (sodium has 11 protons)
c. In a neutral atom, number of electrons equals number of protons

**Mass Number**

Mass number—the number of protons and neutrons in a particular element

a. Nomenclature: $^{23}\text{Na}$ (sodium has a total of 23 neutrons and protons)
b. Knowing both the atomic number and mass number, you can determine the number of neutrons ($23 - 11 = 12$)
c. Mass number is an approximate measure of the mass of an atom since proton and neutron mass are approximately 1 dalton (Na has an atomic mass of approximately 23 daltons)
d. Number of protons in an atom is constant; number of neutrons varies (see below)

**Atomic Weight**

Atomic Weight—the weighed mean of the masses of an element’s constituent isotopes

**Isotope**

Isotope—the of an element that have the same atomic number but different mass number

a. Isotopes of an element have the same number of protons and different number of neutron
b. In natural conditions, elements occur as mixtures of isotopes
c. Different isotopes of the same element react chemically in the same way
d. Isotopes can be radioactive

**Radioactive Isotope**

Radioactive Isotope—an unstable isotope that spontaneously emits nuclear particles and/or energy (i.e., decays)

a. Loss of nuclear particles may transform one element into another
b. Isotopes have characteristic decay times—*half-life*.
c. Half-life—time for 50% of radioactive atoms in a sample to decay

**Biological Applications of Radioactive Isotopes**

1. Dating of geological strata and fossils
   a. Decay occurs at a fixed rate
   b. Compare ratio of radioactive and stable isotopes in fossils with the ratio in living organisms (e.g., $^{14}\text{C}$ and $^{12}\text{C}$)
2. Radioactive tracers
a. Course of biological reactions can be monitored or location of particular substances can be determined
b. Particular substances have unique chemistry (e.g., DNA) allowing specific labeling (e.g., $^{32}$P)
c. Treatment of cancer—cobalt
d. Specific technologies employ radioisotopes (e.g., PET)

Definitions:
Energy—ability to do work
Potential energy—energy stored by matter due to its position or location
a. Matter tends to move to the lowest state of potential energy
b. Exists in discrete units—quanta (i.e., not infinitely divisible)

Energy Levels of Electrons
Electrons are negatively charged particles that orbit around the nucleus.

a. Equal in mass and charge
b. Involved in chemical reactions (only subatomic particle)
c. Have potential energy
d. Electrons have fixed potential energy states called electron shells
e. Electrons with the lowest potential energy are closest to the nucleus
f. Electrons with the highest potential energy are farthest from the nucleus
g. Electrons are not fixed in a single orbit but can move from one energy level to another, thereby gaining or losing potential energy.

Electron Orbitals
Orbitals are three-dimensional space where electrons will be found 90% of the time.
a. Statistical model—viewed as a three-dimensional probability cloud
b. Number of electrons in any given cloud can not exceed two
c. Energy levels
   i. 1<sup>st</sup> energy level
      (a) One spherical s orbital (1s orbital)
      (b) 1<sup>st</sup> energy level can have a maximum of two electrons
   ii. 2<sup>nd</sup> energy level
      (a) Maximum of eight electrons
      (b) One spherical s orbital (2s orbital)
      (c) Three dumbbell-shaped p orbitals each oriented at right angles to the other two (2p<sub>x</sub>, 2p<sub>y</sub>, 2p<sub>z</sub> orbitals)
   iii. Higher energy levels
      (a) Contain s and p orbitals
      (b) Additional orbitals with more complex shapes
d. Movement between energy levels requires gain (up) or loss (down) of energy.

Electron Configuration Determines an Atom’s Chemical Properties.
Electron configuration is the distribution of electrons in an atom’s electron shells.

Elements (first 18; first three rows) of the periodic table:
a. Ordered by atomic number
b. Outermost shell of these atoms never has more than four orbitals (2s and 2p<sub>x</sub>, 2p<sub>y</sub>, 2p<sub>z</sub>) or eight electrons
c. Electrons assume the lowest potential energy state first (i.e., lower electron shells before higher; electrons 1 and 2 into 1s)
d. Electrons are added to p orbitals singly, then paired
e. Outer shells may only be partially filled
   i. Example: $^8$O has eight electrons
Valence Electrons
Valence electrons are the electrons in the outermost shell. The chemical properties of an atom depend on these electrons.

- Octet rule—valence shell is complete when it contains eight electrons (except H and He).
- An atom with a complete valence shell is inert or unreactive.
- Noble gases (e.g., helium argon, and neon) are inert in their elemental state.
- Atoms without complete valence shells are chemically reactive and can form chemical bonds until eight electrons fill the valence shell.
-Atoms with the same number of electrons exhibit similar chemical behavior.
- In general, valence electrons are responsible for an atom’s binding capacity.

Molecules
Atoms combine by chemical binding to form molecules. Atoms with incomplete valence shells tend to fill those shells by interacting with other atoms. The consequence of this interaction may be the formation of a chemical bond.

Terms:
- Chemical bond—attraction that holds a molecule together.
- Molecule—two or more atoms held together by chemical bonds.

Types of Chemical Bonds
1. Covalent bond—chemical bond between atoms formed by sharing a pair of valence electrons.
   - Strong chemical bond (e.g., H₂, O₂, N₂)

Structural versus molecular formulas
- Structural—represents the atoms and the bonding within a molecule (e.g., H–H); the line represents the bond.
- Molecular—number and type of atoms but no indication of arrangement of bonds (e.g., CH₄).

- Single covalent bonds—bond formed by sharing a single pair of valence electrons.
- Double covalent bonds—bond formed by sharing two pairs of valence electrons.
- Triple covalent bonds—bond formed by sharing three pairs of valence electrons.
- Double and triple bonds are rigid and do not allow rotation.

Terms:
- Valence—number of covalent bonds that must be formed to complete the outer electron shell (i.e., bonding capacity of an atom).

Examples:

<table>
<thead>
<tr>
<th></th>
<th>Valence</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1</td>
</tr>
<tr>
<td>O</td>
<td>2</td>
</tr>
<tr>
<td>N</td>
<td>3</td>
</tr>
</tbody>
</table>
Compound—a pure substance composed of two or more elements combined in a fixed ratio (e.g., H₂O, CH₄)

Electronegativity—a atom’s ability to attract and hold electrons
   a. The greater an atom’s electronegativity, the more strongly it attracts shared electrons
   b. Scale:
      | Electronegativity |
      |--------|
      | O      | 3.5   |
      | N      | 3.0   |
      | S      | 2.5   |
      | C      | 2.5   |
      | P      | 2.1   |

2. Types of Covalent Bonds—Nonpolar and polar

Nonpolar covalent bond—a bond formed by equal sharing of electrons between atoms
   a. Occurs when the electronegativity of two atoms is roughly equal (e.g., CH₄)
   b. Molecule made of one element usually have nonpolar covalent bonds

Polar covalent bond—a bond formed by an unequal sharing of electrons between atoms
   a. Occurs when the electronegativity of two atoms is different
   b. Shared electrons spend more time in proximity to the atom with higher electronegativity
   c. For H₂O, the oxygen is strongly electronegative. The electrons spend more time near the oxygen atom, creating localized charges. The hydrogen have a partial positive charge and the oxygen a partial negative charge (see next lecture).

3. Ionic Bonds
Terms:
Ion—a charged atom or molecule
Anion—an atom that has gained one or more electrons from another atom and has become negatively charged; a negatively charged ion
Cation—an atom that has lost one or more electrons and has become positively charged; a positively charged ion
Ionic bond—a bond formed by the electrostatic attraction made possible by the complete transfer of an electron from a donor atom to an acceptor
   a. Attraction results from the strength of electronegativity
   b. Ionic bonds are strong in crystal form but weak in water
   c. Ionic compounds are called salts

Electronegativity continuum:
LOW Nonpolar covalent→Polar covalent→Ionic HIGH

4. Weak Chemical Bonds
   a. Types:
i. Hydrogen bonds
ii. Van der Waals forces
iii. Hydrophobic interactions
b. Temporary associations—bonds that form briefly and reversibly (e.g., receptor binding)
c. Occur between different molecules or different parts of a single large molecule
d. Help stabilize tertiary structure of large molecules

Hydrogen Bonds
Hydrogen bonds are formed by the charge attraction when a hydrogen atom covalently bonded to a one electronegative atom is attracted to another electronegative atom.
   a. 20X easier to break than covalent bond
   b. Charge attraction between oppositely charged parts of polar molecules

Van der Waal Interactions
Occur when molecules are very close and asymetries result in neighboring electron clouds.

Biological Function is Related to the Shape of the Molecule
Molecules have characteristic sizes and shapes. For many molecules, biological function depends on molecular shape. Receptor binding generally involves spatial conformation.
Molecular shape is determined by nature of bonds.
   a. Molecules with two atoms can only be linear
   b. Molecules with more than two atoms can have complex shapes
   c. Bond formation results in rearrangement of valence shells.

Bonds are Created and Broken by Chemical Reactions
Chemical reaction is the process by which the composition of matter is changed by making or breaking chemical bonds.
   a. Reactants into products
   b. Matter is conserved; reactant atoms are only rearranged
   c. Most reactions are reversible
   d. Reaction rate is affected by the relative concentration of reactants and products
   e. Chemical reactions proceed in both directions
   f. When chemical equilibrium is reached, the rate of forward and reverse reaction are equal and the relative concentration of reactants and products stop changing

II. Inorganic Chemistry: Water
Correlational support for the view that water contributes to the fitness of the environment to support life:
1. Life is likely to evolved in water
2. Living cells are comprised of 70-95% water
3. Water covers 75% of the earth’s surface
4. Water naturally occurs in all three physical states of matter (i.e., solid, liquid and gas)

Water’s properties result from its unique structure and molecular interactions.
Polarity

*Water is a polar molecule* (see last lecture).
1. Water has four valence orbitals that are aligned in a tetrahedron.
2. Two corners of the tetrahedron are orbitals with unshared pairs of electrons and weak negative charge.
3. Two of the corners are occupied with \( H \) atoms in polar covalent bonds with \( O \).
4. Because \( O \) is strongly electronegative, shared electrons with \( H \) spend more time around the \( O \) atom, creating a weak positive charge near the \( H \)'s.

*Due to hydrogen bonding, water has a high level of structural organization.*

1. Individual polar water molecules are held together by hydrogen bonds.
2. The localized charges on one water molecule are attracted to opposite charges on other water molecules.
3. Since there are four orbitals carrying weak charges, water molecules can form hydrogen bonds with four neighboring water molecules.

*From water’s polarity and capacity to form hydrogen bonds, a number of properties emerge:

**Cohesive Behavior**—cohesion results from the tendency of a substance (e.g., water) to be held together by hydrogen bonds.
1. Though transient, the number of hydrogen bonds at any given time gives water more structure than other liquids.*
2. Hydrogen bonding holds water column together in plants, counteracting the downward pull of gravity—water adheres to vessel walls.

3. **Surface tension**—measure of the difficulty to stretch or break the surface of a liquid
   a. At the air/water interface, surface water molecules are hydrogen bonded to each other and to other molecules below the surface
   b. Relative to other liquids, water has a greater surface tension
   c. Surface tension causes water to bead (i.e., the shape with smallest area to volume ratio the permits hydrogen bonding)

**High Specific Heat**

Terms:
- **Kinetic energy**—energy of motion
- **Heat**—total kinetic energy due to molecular motion in a body of matter
- **Temperature**—measure of heat intensity due to the average kinetic energy of molecules in a body of matter

<table>
<thead>
<tr>
<th></th>
<th>Celsius (°C)</th>
<th>Fahrenheit (°F)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water Boils</td>
<td>100</td>
<td>212</td>
</tr>
<tr>
<td>Human Body Temperature</td>
<td>37</td>
<td>98.6</td>
</tr>
<tr>
<td>Room Temperature</td>
<td>23</td>
<td>72</td>
</tr>
<tr>
<td>Water Freezes</td>
<td>0</td>
<td>32</td>
</tr>
</tbody>
</table>

Conversion Between Celsius and Fahrenheit:

\[ ^\circ C = \frac{5(\circ F - 32)}{9} \]
\[ ^\circ F = \frac{9^\circ C}{5} + 32 \]
\[ ^\circ K = ^\circ C = 273 \]

- **Calorie (cal)**—amount of heat required to raise one gram of water one degree Celsius (amount of heat released when one gram of water cools one degree Celsius)
- **Kilocalorie (kcal)**—amount of heat required to raise one kilogram of water one degree Celsius (amount of heat released when one kilogram of water cools one degree Celsius)
- **Specific heat**—amount of heat that must be absorbed or lost for one gram of a substance to change its temperature one degree Celsius
- **Specific heat of water**—1 cal/g°C

Water has a high specific heat (i.e., it resists temperature changes when it absorbs or releases heat).

Keep in mind that heat is a measure of energy resulting from molecular movement. Since hydrogen bonding limits water’s capacity to move, these bonds must be disrupted to permit water molecules to move faster (increase temperature). To break (form) a bond, energy must be absorbed (released). As a result of hydrogen bonding among water molecules, it takes a relatively large heat loss or gain for each 1°C change in temperature. **Biological Implications**

The combination of water’s high heat capacity and its prevalence in the environment act as a heat sink.

1. Energy absorbed during the day and summer only moderately warms large bodies of water.
2. This energy is released during the night and winter.

Water’s heat capacity also contributes to:
1. Relative stability of the temperature of the marine environment
2. Milder climate of coastal areas
3. Regulation of body temperature

**Heat of Vaporization**

Terms:
Vaporization—transformation of a liquid into a gas
Heat of vaporization—quantity of heat required for 1 gram of liquid to be converted into gas
Evaporative cooling—cooling of a liquid’s surface when liquid evaporated

1. Molecules with sufficient kinetic energy can overcome the mutual attraction of molecules in a liquid and escape into the air.
2. For water to evaporate, energy is required to break hydrogen bonds.
3. Relative to other liquids, water has a high heat of vaporization (540 cal/g or 2260 J/g; Joule=0.239 cal)
4. Molecules at a liquid’s surface with the greatest kinetic energy are most likely to escape into gaseous form. The average kinetic energy (i.e., temperature) of the remaining molecules is lower.

**Biological Implications**
1. Evaporative cooling moderates the Earth’s climate—evaporation of tropical surface water helps to dissipate absorbed solar energy
2. As moist tropical air (i.e., air with molecules of high kinetic energy) moves towards colder climates, heat is released as water condenses into rain (liquid)
3. Some animals can regulate body temperature by evaporative cooling

**Density of Solid Water**

Water is less dense in solid than liquid form. This characteristic is the result hydrogen bonding.
1. As water cools (i.e., loses kinetic energy) it contracts.
2. Water is densest at 4°C.
3. From 4°C to freezing (0°C), water expands, becomes less dense and floats.
4. The densest water can be found at the bottom of a body of water, the least dense at the surface, and water with intermediate density lying in the middle.
5. Keeping in mind that temperature is a measure of kinetic energy (i.e., measure of heat intensity due to the average kinetic energy of molecules in a body of matter), as kinetic energy is reduced, capacity to move is also reduced (i.e., direct relationship). At warmer temperatures, water’s kinetic energy permits hydrogen bonds to be broken. At near freezing temperatures, water no longer has the necessary kinetic energy to break hydrogen bonds.
6. When each water molecule has four hydrogen bonds, a crystalline lattice is formed. As a crystal, each molecule is maximally separated. (Density is a measure of the number of molecules within a given physical space. For water,
molecules are maximally separated in the solid form and therefore it is least dense.)

**Biological Implications**
1. Deep bodies of water do not freeze solid from the bottom up.
2. The process of freezing releases energy to the water lying below the solid water. This in effect insulates (warms) the liquid water near the newly forming ice.
3. Dampens seasonal changes. Energy associated with the formation or breaking of hydrogen bonds, moderates temperature of the environment during seasonal changes.

**Water as a Solvent**
Terms:
Solution—a liquid consisting of a completely homogeneous mixture of two or more substances
Solvent—dissolving agent of a solution
Solute—substance dissolved in a solution
Aqueous solution—solution in which is the solvent
Hydrophobic—“water fearing”; non-water soluble compounds; generally nonpolar compounds with symmetric charge distribution
Hydrophilic—“water loving”; water soluble compounds
1. Ionic compounds—held together by ionic bonds formed by electrostatic attraction following the complete transfer of an electron from a donor atom
2. In polar solvents, the charged region of the polar molecule has an electrical attraction to charged ions
3. Water surrounds he individual charged ions, shielding them from each other
4. Polar solvents also have affinity for oppositely charged regions of other polar molecules

**Quantitative Properties of Aqueous Solutions:**
1. Solute Concentration
2. pH
Terms:
Molecular weight—sum of the weight of all atoms in a molecule
Mole—amount of a substance that has a mass in grams numerically equivalent to its molecular weight in daltons

E.g.: Sucrose (C\textsubscript{12}H\textsubscript{22}O\textsubscript{11})

<table>
<thead>
<tr>
<th>Atomic Weight</th>
<th>Number of Atoms</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>12 dalton 12</td>
</tr>
<tr>
<td>H</td>
<td>1 dalton 22</td>
</tr>
<tr>
<td>O</td>
<td>16 dalton 11</td>
</tr>
</tbody>
</table>

| Molecular Weight | 342 dalton |

Therefore, one mole of sucrose is equal to 342 grams.
Molarity—number of moles of solute per liter of solution
To make a 1M solution of sucrose, weigh out 342 grams of sucrose and add water until the total volume equals 1 liter.

**Why do we use moles?**
1. From a practical standpoint, grams are more convenient (1 gram = 1.67 x 10^{24} dal)
2. One mole of any substance will have the same number of molecules (6.02 x 10^{23})
3. Substances can be combined in fixed ratios based on number of molecules

**Dissociation of Water**
An individual water molecule is comprised of two hydrogen atoms in covalent bonds with an oxygen atom. The hydrogen atom is usually shared in a hydrogen bond between two water molecules. Occasionally, the hydrogen atom will shift from the oxygen atom to which it is covalently bound to the unshared orbitals of the oxygen to which it is hydrogen bonded. Details:
1. Only a hydrogen ion (proton with a +1 charge) is transferred, not the electron
2. The transferred proton binds to an unshared orbital of the second water molecule forming a hydronium ion (H_3O^+)
3. The water molecule that lost a proton has a net negative charge and is called a hydroxide ion (OH^-)
4. The ionization reaction is reversible and exists in an equilibrium where most of the H_2O is not ionized.

Reaction:
H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-

By convention, this reaction is expressed as the dissociation of water into H^+ and OH^-:
H_2O \rightleftharpoons H^+ + OH^-

**[H^+] and pH**

**Acids and Bases**
At equilibrium in pure water at 25°C:
1. The number of H^+ ions is equal to the number of OH^- ions ([H^+] = [OH^-])
2. Under these conditions there are very few water molecules that are dissociated—1 out of 5.54 x 10^8. The concentration of hydronium and hydroxyl ions is 1 gram of ions per each 1.0 x 10^7 liters or 1 x 10^{-7} M.

**Acids**
By definition, an acid is a substance that increases the relative [H^+] of a solution. Acids combine with OH^- to form H_2O.

**Bases**
By definition, a base is a substance that decreases the relative [H^+] of a solution; alternately, bases increase [OH^-].

E.g.:
Acid: HCl \rightarrow H^+ + Cl^-
Base
Direct Reduction: NH_3^+ + H^+ \rightleftharpoons NH_4^+
Indirect Reduction: $\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$

$\text{OH}^- + \text{H}^+ \Leftrightarrow \text{H}_2\text{O}$

<table>
<thead>
<tr>
<th>Solution</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Neutral</td>
<td>$[\text{H}^+] = [\text{OH}^-]$</td>
</tr>
<tr>
<td>Acidic</td>
<td>$[\text{H}^+] &gt; [\text{OH}^-]$</td>
</tr>
<tr>
<td>Basic</td>
<td>$[\text{H}^+] &lt; [\text{OH}^-]$</td>
</tr>
</tbody>
</table>

*Strong Acids and Bases*
1. Completely dissociate in water
2. Single arrow indicates a complete dissociation
   $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$

*Weak Acids and Bases*
1. Partially and reversibly dissociate
2. Indicated by a double arrow

*Measuring acidity: pH scale*
1. In any aqueous solution $[\text{H}^+][\text{OH}^-] = 1 \times 10^{-14}$
   a. Neutral: $[\text{H}^+]$ and $[\text{OH}^-]$ = $1 \times 10^{-7}$
   b. In an acidic solution where $[\text{H}^+] = 1 \times 10^{-5}$, $[\text{OH}^-] = 1 \times 10^{-9}$
   c. In a basic solution where $[\text{H}^+] = 1 \times 10^{-9}$, $[\text{OH}^-] = 1 \times 10^{-5}$
2. The pH scale measures degree of acidity.
   a. It ranges from 0 to 14.
   b. pH = $-\log_{10}$ of the $[\text{H}^+]$ expressed in moles per liter
   c. pH of 7 is a neutral solution
   d. pH < 7 is an acidic solution
   e. pH > 7 is a basic solution
   f. Each 1 unit change in pH represents a 10-fold increase in $[\text{H}^+]$

How much greater is the $[\text{H}^+]$ in a solution with a pH of 2 ($-\log_{10} 1 \times 10^{-2}$) than one with a pH of 6 ($-\log_{10} 1 \times 10^{-6}$)? Answer: 10,000 (1/100 vs. 1/1,000,000)

3. Most biological systems have pH’s ranging from 6 to 8 (a fairly narrow range).
   Exception: pH of stomach acid is 1.5.

*Buffers*
Buffers are substances that minimize changes in pH. In biological systems with narrow pH ranges, buffers are very important.
1. Buffers are combinations of H$^+$-donor and H$^+$-acceptor forms in a solution of weak acids or bases.
2. Function by accepting H$^+$ ions from solution when in excess and by donating H$^+$ ions to the solution when they have been depleted.

$\text{H}_2\text{CO}_3 \leftrightarrow \text{HCO}_3^- + \text{H}^+$
1. $\text{H}_2\text{CO}_3$ is a weak acid and acts as a H$^+$-donor
2. \(\text{HCO}_3^-\) is a weak base and acts as a \(\text{H}^+\)-acceptor
3. In response to a rise in pH, reaction shifts to the right with the production of hydrogen ion (\(\text{H}^+\))
4. In response to a decrease in pH, reaction shifts to the left

E.g.:
\[
\text{HCl} + \text{Na HCO}_3 \rightarrow \text{H}_2\text{CO}_3 + \text{NaCl} \\
\text{NaOH} + \text{H}_2\text{CO}_3 \rightarrow \text{Na HCO}_3 + \text{H}_2\text{O}
\]

**Acid Precipitation**
Rain, snow or fog more strongly acidic than pH 5.6.
Primarily due to sulfur oxides and nitrogen oxides produced by combustion of fossil fuels. In the atmosphere, these oxides react with water in the air to form acids.

Acid precipitation:
2. Lowers pH of lakes and ponds

### III. Organic Chemistry—a branch of chemistry that specializes in the study of carbon compounds

**Terms:**
Organic molecules—molecules containing carbon
Vitalism—belief the life force is not governed by chemical or physical laws; prior to the 19th century, organic compounds could not be artificially synthesized
Mechanism—natural phenomena are governed by chemical and physical laws

1. Mainstream biological favors this view.

**Carbon**
1. Atomic number of 6 (6 protons and 6 electrons)
2. Carbon has 4 valence electrons.
3. Does not normally form ionic bonds
4. Completes outer shell by forming 4 covalent bonds; tetravalent electron configuration

**Tetravalent electron configuration**
1. Permits the formation of large, complex molecules
2. Carbon is oriented in the center from which the molecule branches in four directions
3. Permits carbon to covalently bind with many different elements, most commonly:

<table>
<thead>
<tr>
<th></th>
<th>Hydrogen</th>
<th>Oxygen</th>
<th>Nitrogen</th>
<th>Carbon</th>
</tr>
</thead>
<tbody>
<tr>
<td>Valence</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>4</td>
</tr>
<tr>
<td>Atomic #</td>
<td>1</td>
<td>8</td>
<td>7</td>
<td>6</td>
</tr>
</tbody>
</table>

4. Determines organic molecule’s three dimensional shape (e.g., four single covalent bonds creates a tetrahedron)

**Carbon Skeletons**
Organic molecules often include long chains of carbons covalently linked together. These carbon skeletons may vary in:

1. **Length**
2. **Shape**
3. **Number and location of double bonds**
4. **Nature of other covalently bound elements**

*Hydrocarbons*—contain only carbon and hydrogen

1. Major component of fossil fuels
2. Since C-H and C-C bonds are nonpolar, hydrocarbons are hydrophobic

**Isomers**

Isomer—compounds sharing the same molecular formula but with different structures and therefore have different properties

Types of isomers:

1. **Structural isomers**—isomers that differ in the covalent arrangement of their atoms
   a. Number of possible isomers increase with carbon skeleton size
   b. May also differ in location of double bonds
2. **Geometric isomers**—share the same covalent partners but differ in spatial arrangements
   a. Double bonds prevent free rotational around central axis
   b. Result in subtle differences in biological activity
3. **Enantiomers**—isomer with mirror images of each other (l-isomer and d-isomer)
   a. Possible when four different atoms or groups of atoms are bonded to the same carbon (i.e., asymmetric carbon)
   b. There are two different possible spatial arrangements of the four groups around the asymmetric carbon
   c. Usually, only one form is biologically active.

**Functional Groups**

There are a number of small groups of atoms that are frequently bonded to the carbon skeleton of organic molecules. These groups are known as *functional groups*.

Functional groups:

1. Have specific chemical and physical properties
2. Are regions of organic molecules that are commonly chemically reactive
3. Exhibit common properties independent of which organic molecule it is bonded to
4. Their number and arrangement determine the unique chemical properties of organic molecules

*Hydroxyl Group*

Hydroxyl group consists of a hydrogen atom bonded to an oxygen atom, which in turn is bonded to carbon (–OH).

Properties:

1. Polar; O-H is a polar covalent bond
2. Water soluble; water can form a hydrogen bond to the polar hydroxyl group
3. Organic compounds with hydroxyl groups are called *alcohols*
Carbonyl Group
Carbonyl group consists of a carbon atom double bonded to an oxygen atom (-CO).
1. Polar; the oxygen atom can be involved in hydrogen bonds
2. Water soluble
3. Functional group found in sugars
4. If the carbonyl group is off the end of the carbon skeleton, it is referred to as an aldehyde.
5. If the carbonyl group is at the terminal end of the carbon skeleton, it is referred to as a ketone.

Carboxyl Group
Carboxyl group consists of a carbon double bonded to one oxygen atom and single bonded to the oxygen of a hydroxyl group (-COOH).
1. Polar and water soluble; covalent bond between oxygen and hydrogen is so polar that the hydrogen reversibly dissociates as H⁺. This polarity results from the combined effect of the two electronegative oxygen atoms bonded to the same carbon.
2. Since it donates protons, it is an acid. Compounds with this functional group are called carboxylic acids.

Amino Group
Amino group consists of a nitrogen atom bonded to two hydrogen atoms and to the carbon skeleton (-NH₂).
1. Polar and water soluble
2. Acts as a weak base; the unshared pair of electrons on the nitrogen can accept a proton, giving the amino group a +1 charge.
3. Compounds with this functional group are called amines.

Sulphydryl Group
Sulphydryl group consists of an atom of sulfur bonded to an atom of hydrogen (-SH).
1. Participates in disulfide bridges that stabilize the tertiary structure of macromolecules
2. Compounds with this functional group are called thiols.

Phosphate Group
Phosphate group is the dissociated form of phosphoric acid (H₃PO₄).
1. Polar and water soluble
2. Dissociation of two hydrogen atoms leaves the phosphate group with a negative charge (-2)
3. Important for cellular energy storage

Methyl Group
Methyl group consists of carbon bonded to three hydrogen atoms (-CH₃).
1. Nonpolar and hydrophobic
2. Contribute to the tertiary structure of macromolecules