

COURSE OUTLINE: INTRODUCTION TO THE CHEMISTRY OF AMINO ACIDS, PROTEINS, AND BIOMOLECULES

1. Introduction to Amino Acids

Definition and Structure of Amino Acids

General structure of amino acids

α -carbon, amino group, carboxyl group, R-group (side chains)

Properties of Amino Acids

Physical properties (solubility, melting point, etc.)

Chemical properties (isoelectric point, acid-base behavior)

Reactions of Amino Acids

Peptide bond formation

Reactions with acids and bases, oxidation, decarboxylation

Biological Functions of Amino Acids

Role in protein synthesis

Precursor to neurotransmitters, hormones, and other biomolecules

2. Classification of Amino Acids

Based on Polarity

Non-polar amino acids (e.g., glycine, alanine)

Polar amino acids (e.g., serine, threonine)

Based on Charge

Neutral amino acids

Acidic amino acids (e.g., aspartic acid, glutamic acid)

Basic amino acids (e.g., lysine, arginine)

Based on Nutritional Value

Essential amino acids (e.g., leucine, valine)

Non-essential amino acids (e.g., alanine, asparagine)

3. Peptides

Definition of Peptides

Difference between peptides and proteins

Peptide Bond Formation

Biological Importance of Peptides

Hormonal and enzymatic functions

4. Introduction to Proteins

Definition and Classification of Proteins

Simple and conjugated proteins

Fibrous and globular proteins

Biological Functions of Proteins

Structural, enzymatic, transport, and immune roles

Methods of Protein Isolation and Purification

Centrifugation, chromatography, electrophoresis

Identification Methods

Spectrophotometry, SDS-PAGE, amino acid sequencing

5. Structure of Proteins

Primary Structure

Amino acid sequence, peptide bond formation

Secondary Structure

Alpha-helices, beta-sheets, hydrogen bonding

Tertiary Structure

3D folding, hydrophobic interactions, disulfide bonds

Quaternary Structure

Multiple polypeptide chains, subunit arrangement

6. Protein and Amino Acid Tests

Basic Tests for Proteins

Biuret test, Xanthoproteic test, Ninhydrin test

Basic Tests for Amino Acids

Thin-layer chromatography, paper chromatography

7. Introduction to Carbohydrates

Definition and Classification

Monosaccharides, disaccharides, oligosaccharides, polysaccharides

Properties and Biological Functions

Energy storage, structural roles in plants and animals

8. Introduction to Lipids

Definition and Classification

Simple, compound, and derived lipids

Properties and Biological Functions

Energy storage, insulation, cell membrane structure

9. Introduction to Nucleic Acids

Definition and Types

DNA, RNA

Nomenclature of Nucleosides and Nucleotides

Purines, pyrimidines, ribose and deoxyribose sugars, phosphate groups

Hydrolysis of Nucleic Acids

Effects of acid and alkali on nucleic acid hydrolysis

10. Applications in Biochemistry

Real-life Applications of Amino Acids and Proteins

Pharmaceuticals, nutrition, disease mechanisms

Clinical Relevance

Protein denaturation, genetic mutations affecting protein function

11. Conclusion and Review

Summary of Key Concepts

Review of Course Objectives

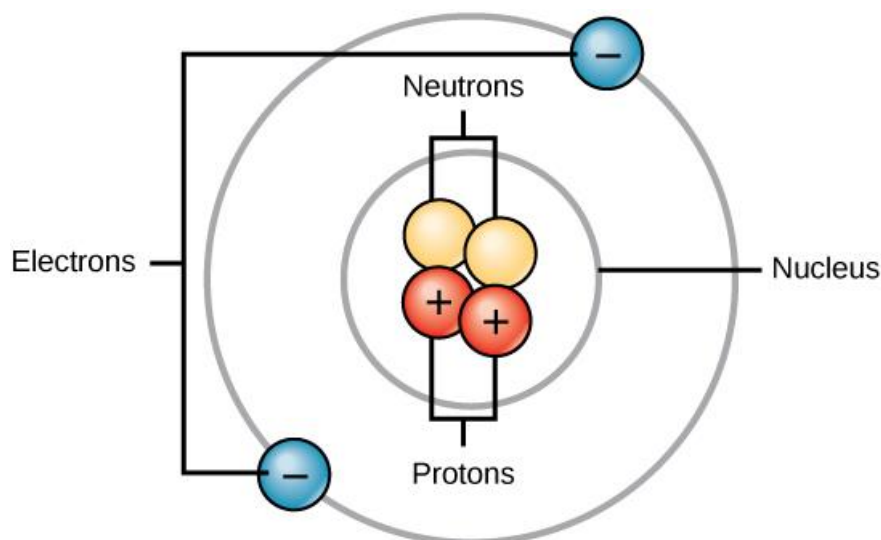
MODULE 1 INTRODUCTION TO BIOCHEMISTRY

Biochemistry is the branch of science that explores the chemical processes and substances that occur within living organisms. It combines principles of both biology and chemistry to understand the molecular mechanisms that support life. By studying biochemical processes, researchers can comprehend how cells communicate, produce energy, and maintain homeostasis. Biochemistry is the study of the molecular and chemical reactions that facilitate life processes. It includes the study of molecules like proteins, lipids, carbohydrates, and nucleic acids, as well as their interactions in metabolic pathways. Biochemistry provides the molecular basis for understanding cellular processes. From DNA replication to metabolism, biochemistry explains how living organisms grow, reproduce, and adapt to their environment. It bridges the gap between biology and chemistry, showing how chemical changes in cells affect their function. Many diseases arise from biochemical imbalances or disruptions. For example, diabetes is caused by irregularities in glucose metabolism, while cancer can involve uncontrolled cell division due to mutations in biochemical pathways. Biochemistry helps develop treatments by targeting these molecular pathways with drugs or gene therapy.

THE CHEMICAL FOUNDATION OF LIFE

ATOMS

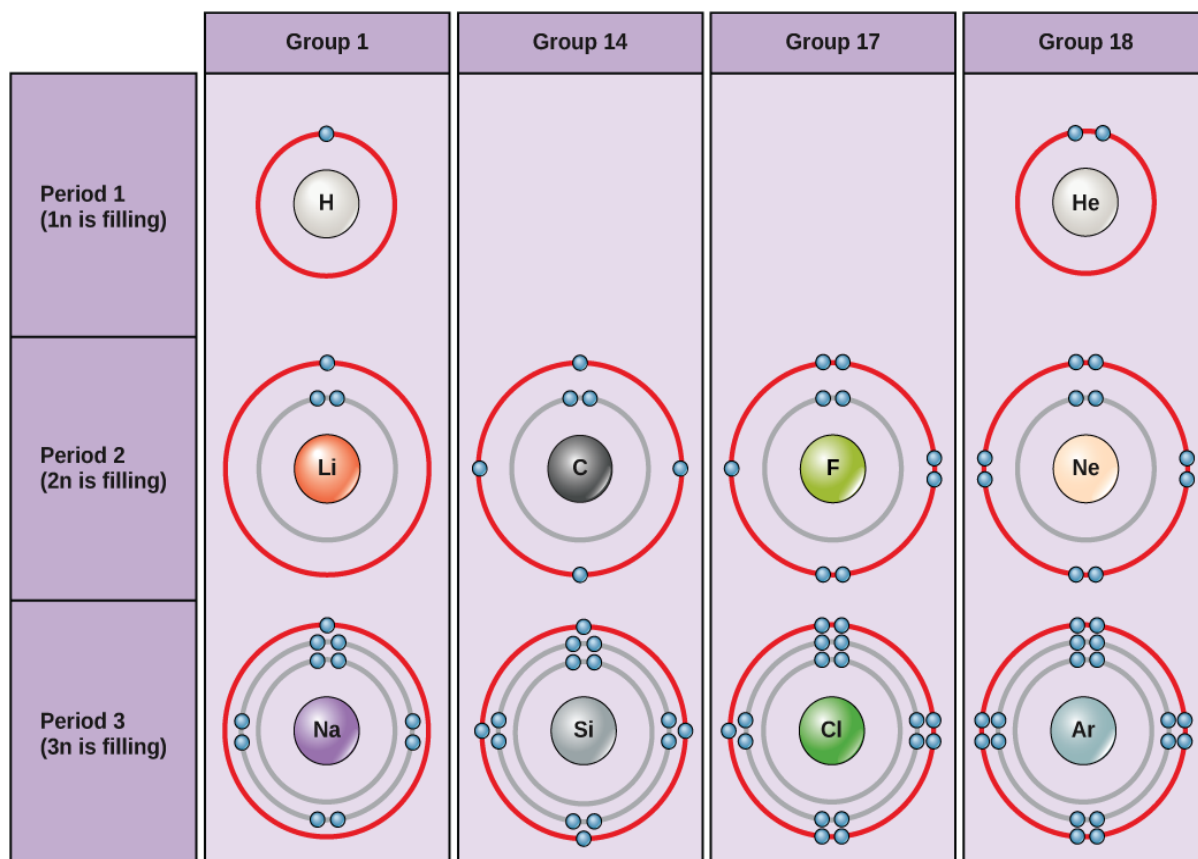
An atom is the smallest component of an element that retains all of the chemical properties of that element. For example, one hydrogen atom has all of the properties of the element hydrogen, such as it exists as a gas at room temperature, and it bonds with oxygen to create a water molecule. Hydrogen atoms cannot be broken down into anything smaller while still retaining the properties of hydrogen. If a hydrogen atom were broken down into subatomic particles, it would no longer have the properties of hydrogen. At the most basic level, all organisms are made of a combination of elements. They contain atoms that combine together to form molecules. In multicellular organisms, such as animals, molecules can interact to form cells that combine to form tissues, which make up organs. These combinations continue until entire multicellular organisms are formed. All atoms contain protons, electrons, and neutrons. The most common isotope of hydrogen (H) is the only exception and is made of one proton and one electron with no neutrons. A **proton** is a positively charged particle that resides in the **nucleus** (the core of the atom) of an atom and has a mass of 1 and a charge of +1. An **electron** is a negatively charged particle that travels in the space around the nucleus. In other words, it resides outside of the nucleus. It has a negligible mass and has a charge of -1.



Neutrons, like protons, reside in the nucleus of an atom. They have a mass of 1 and no charge. The positive (protons) and negative (electrons) charges balance each other in a neutral atom, which has a net zero charge. Because protons and neutrons each have a mass of 1, the mass of an atom is equal to the number of protons and neutrons of that atom. The number of electrons does not factor into the overall mass, because their mass is so small. As stated earlier, each element has its own unique properties. Each contains a different number of protons and neutrons, giving it its own atomic number and mass number. The **atomic number** of an element is equal to the number of protons that element contains. The **mass number**, or atomic mass, is the number of protons plus the number of neutrons of that element. Therefore, it is possible to determine the number of neutrons by subtracting the atomic number from the mass number. These numbers provide information about the elements and how they will react when combined. Different elements have different melting and boiling points, and are in different states (liquid, solid, or gas) at room temperature. They also combine in different ways. Some form specific types of bonds, whereas others do not. How they combine is based on the number of electrons present. Because of these characteristics, the elements are arranged into the **periodic table of elements**, a chart of the elements that includes the atomic number and relative atomic mass of each element. The periodic table also provides key information about the properties of elements often indicated by color-coding. The arrangement of the table also shows how the electrons in each element are organized and provides important details about how atoms will react with each other to form molecules.

Isotopes are different forms of the same element that have the same number of protons, but a different number of neutrons. Some elements, such as carbon, potassium, and uranium, have

naturally occurring isotopes. Carbon-12, the most common isotope of carbon, contains six protons and six neutrons. Therefore, it has a mass number of 12 (six protons and six neutrons) and an atomic number of 6 (which makes it carbon). Carbon-14 contains six protons and eight neutrons. Therefore, it has a mass number of 14 (six protons and eight neutrons) and an atomic number of 6, meaning it is still the element carbon. These two alternate forms of carbon are isotopes. Some isotopes are unstable and will lose protons, other subatomic particles, or energy to form more stable elements. These are called **radioactive isotopes** or radioisotopes.

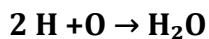


Bohr diagrams indicate how many electrons fill each principal shell. Group 18 elements (helium, neon, and argon are shown) have a full outer, or valence, shell. A full valence shell is the most stable electron configuration. Elements in other groups have partially filled valence shells and gain or lose electrons to achieve a stable electron configuration. An atom may give, take, or share electrons with another atom to achieve a full valence shell, the most stable electron configuration.

CHEMICAL BONDS

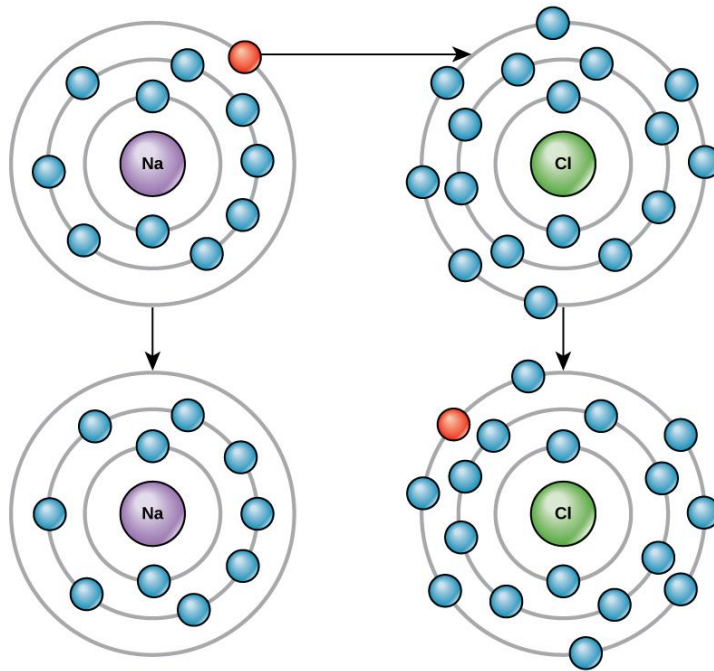
How elements interact with one another depends on how their electrons are arranged and how many openings for electrons exist at the outermost region where electrons are present in an

atom. Electrons exist at energy levels that form shells around the nucleus. The closest shell can hold up to two electrons. The closest shell to the nucleus is always filled first, before any other shell can be filled. Hydrogen has one electron; therefore, it has only one spot occupied within the lowest shell. Helium has two electrons; therefore, it can completely fill the lowest shell with its two electrons.



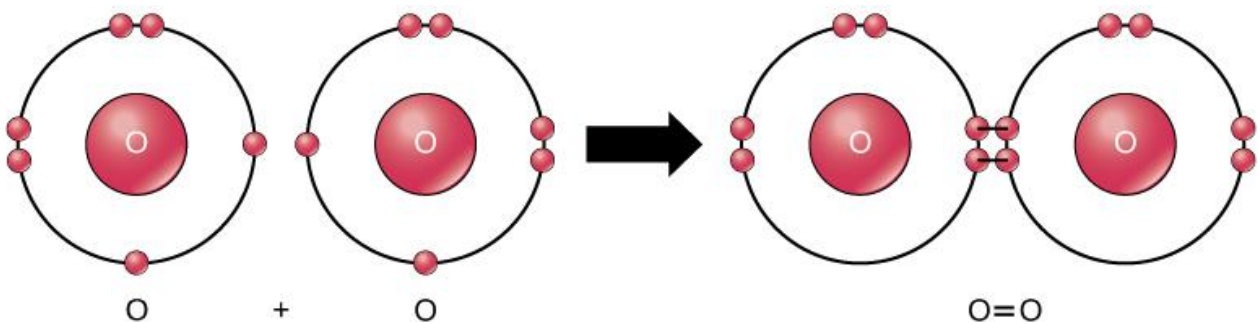
Two or more atoms may bond with each other to form a molecule. When two hydrogens and an oxygen share electron via covalent bonds, a water molecule is formed. Not all elements have enough electrons to fill their outermost shells, but an atom is at its most stable when all of the electron positions in the outermost shell are filled. Because of these vacancies in the outermost shells, we see the formation of **chemical bonds**, or interactions between two or more of the same or different elements that result in the formation of molecules. To achieve greater stability, atoms will tend to completely fill their outer shells and will bond with other elements to accomplish this goal by sharing electrons, accepting electrons from another atom, or donating electrons to another atom. Because the outermost shells of the elements with low atomic numbers (up to calcium, with atomic number 20) can hold eight electrons, this is referred to as the **octet rule**. An element can donate, accept, or share electrons with other elements to fill its outer shell and satisfy the octet rule. When an atom does not contain equal numbers of protons and electrons, it is called an **ion**. Because the number of electrons does not equal the number of protons, each ion has a net charge. Positive ions are formed by losing electrons and are called **cations**. Negative ions are formed by gaining electrons and are called **anions**. For example, sodium only has one electron in its outermost shell. It takes less energy for sodium to donate that one electron than it does to accept seven more electrons to fill the outer shell. If sodium loses an electron, it now has 11 protons and only 10 electrons, leaving it with an overall charge of +1. It is now called a sodium ion.

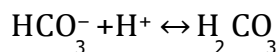
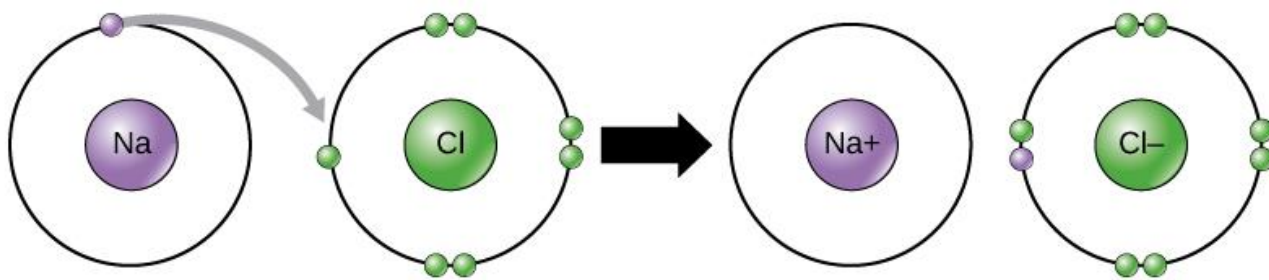
The chlorine atom has seven electrons in its outer shell. Again, it is more energy-efficient for chlorine to gain one electron than to lose seven. Therefore, it tends to gain an electron to create an ion with 17 protons and 18 electrons, giving it a net negative (–1) charge. It is now called a chloride ion. This movement of electrons from one element to another is referred to as **electron transfer**. As sodium atom (Na) only has one electron in its outermost shell, whereas a chlorine atom (Cl) has seven electrons in its outermost shell. A sodium atom will donate its one electron to empty its shell, and a chlorine atom will accept that electron to fill its shell, becoming chloride. Both ions now satisfy the octet rule and have complete outermost shells.



IONIC BONDS

There are four types of bonds or interactions: ionic, covalent, hydrogen bonds, and van der Waals interactions. Ionic and covalent bonds are strong interactions that require a larger energy input to break apart. When an element donates an electron from its outer shell, as in the sodium atom example above, a positive ion is formed. The element accepting the electron is now negatively charged. Because positive and negative charges attract, these ions stay together and form an **ionic bond**, or a bond between ions. The elements bond together with the electron from one element staying predominantly with the other element. When Na^+ and Cl^- ions combine to produce NaCl , an electron from a sodium atom stays with the other seven from the chlorine atom, and the sodium and chloride ions attract each other in a lattice of ions with a net zero charge.





In the formation of an ionic compound, metals lose electrons and nonmetals gain electrons to achieve an octet.

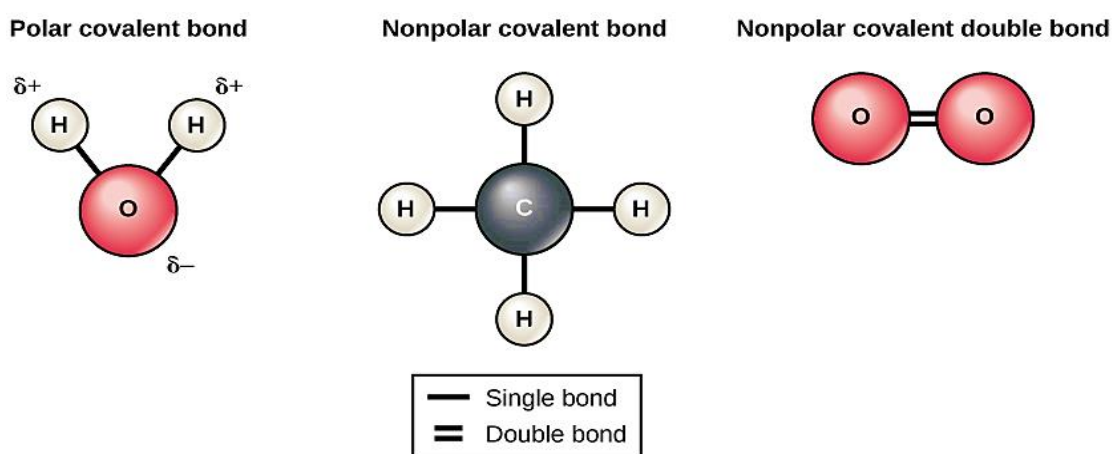
COVALENT BONDS

Another type of strong chemical bond between two or more atoms is a **covalent bond**. These bonds form when an electron is shared between two elements and are the strongest and most common form of chemical bond in living organisms. Covalent bonds form between the elements that make up the biological molecules in our cells. Unlike ionic bonds, covalent bonds do not dissociate in water. The hydrogen and oxygen atoms that combine to form water molecules are bound together by covalent bonds. The electron from the hydrogen atom divides its time between the outer shell of the hydrogen atom and the incomplete outer shell of the oxygen atom. To completely fill the outer shell of an oxygen atom, two electrons from two hydrogen atoms are needed, hence the subscript “2” in H_2O . The electrons are shared between the atoms, dividing their time between them to “fill” the outer shell of each. This sharing is a lower energy state for all of the atoms involved than if they existed without their outer shells filled. There are two types of covalent bonds: polar and nonpolar.

Nonpolar covalent bonds form between two atoms of the same element or between different elements that share the electrons equally. For example, an oxygen atom can bond with another oxygen atom to fill their outer shells. This association is nonpolar because the electrons will be equally distributed between each oxygen atom. Two covalent bonds form between the two oxygen atoms because oxygen requires two shared electrons to fill its outermost shell. Nitrogen atoms will form three covalent bonds (also called triple covalent) between two atoms of nitrogen because each nitrogen atom needs three electrons to fill its outermost shell. Another example of a nonpolar covalent bond is found in the methane (CH_4) molecule. The carbon atom has four electrons in its outermost shell and needs four more to fill it. It gets these four from

four hydrogen atoms, each atom providing one. These elements all share the electrons equally, creating four nonpolar covalent.

In a **polar covalent bond**, the electrons shared by the atoms spend more time closer to one nucleus than to the other nucleus. Because of the unequal distribution of electrons between the different nuclei, a slightly positive (δ^+) or slightly negative (δ^-) charge develops. The covalent bonds between hydrogen and oxygen atoms in water are polar covalent bonds. The shared electrons spend more time near the oxygen nucleus, giving it a small negative charge, than they spend near the hydrogen nuclei, giving these molecules a small positive charge.



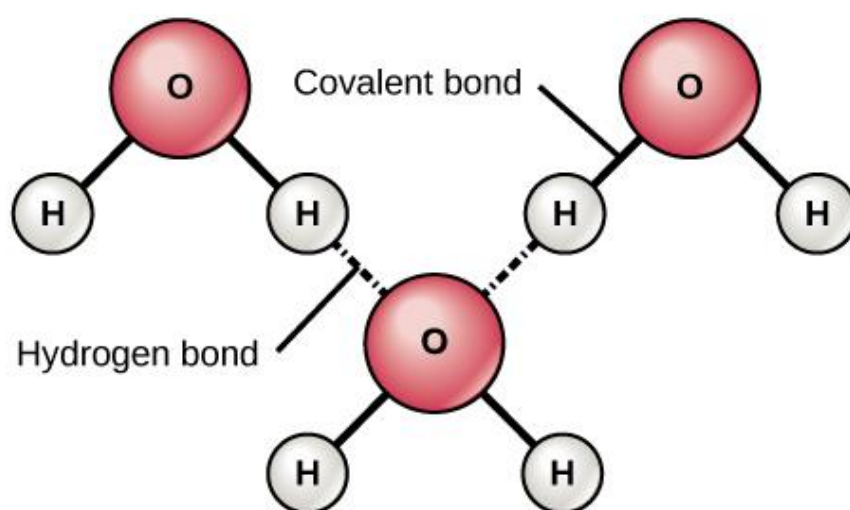
	Bond type	Molecular shape	Molecular type
Water	<p>Polar covalent</p>	<p>Bent</p>	Polar
Methane	<p>Nonpolar covalent</p>	<p>Tetrahedral</p>	Nonpolar
Carbon dioxide	<p>Polar covalent</p>	<p>Linear</p>	Nonpolar

The water molecule (left) depicts a polar bond with a slightly positive charge on the hydrogen atoms and a slightly negative charge on the oxygen. Examples of nonpolar bonds include methane (middle) and oxygen (right).

HYDROGEN BONDS

Ionic and covalent bonds are strong bonds that require considerable energy to break. However, not all bonds between elements are ionic or covalent bonds. Weaker bonds can also form. These are attractions that occur between positive and negative charges that do not require much energy to break. Two weak bonds that occur frequently are hydrogen bonds and van der Waals interactions. These bonds give rise to the unique properties of water and the unique structures of DNA and proteins.

When polar covalent bonds containing a hydrogen atom form, the hydrogen atom in that bond has a slightly positive charge. This is because the shared electron is pulled more strongly toward the other element and away from the hydrogen nucleus. Because the hydrogen atom is slightly positive ($\delta+$), it will be attracted to neighboring negative partial charges ($\delta-$). When this happens, a weak interaction occurs between the $\delta+$ charge of the hydrogen atom of one molecule and the $\delta-$ charge of the other molecule. This interaction is called a **hydrogen bond**. This type of bond is common; for example, the liquid nature of water is caused by the hydrogen bonds between water molecules. Hydrogen bonds give water the unique properties that sustain life. If it were not for hydrogen bonding, water would be a gas rather than a liquid at room temperature.



Hydrogen bonds form between slightly positive ($\delta+$) and slightly negative ($\delta-$) charges of polar covalent molecules, such as water.

Hydrogen bonds can form between different molecules and they do not always have to include a water molecule. Hydrogen atoms in polar bonds within any molecule can form bonds with other adjacent molecules. For example, hydrogen bonds hold together two long strands of DNA to give the DNA molecule its characteristic double-stranded structure. Hydrogen bonds are also responsible for some of the three-dimensional structure of proteins.

van der Waals Interactions


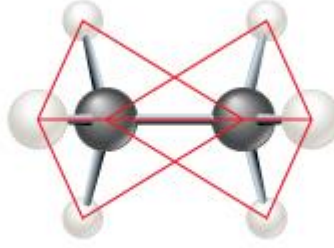
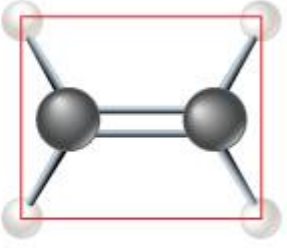
Like hydrogen bonds, **van der Waals interactions** are weak attractions or interactions between molecules. They occur between polar, covalently bound, atoms in different molecules. Some of these weak attractions are caused by temporary partial charges formed when electrons move around a nucleus. These weak interactions between molecules are important in biological systems.

HYDROCARBONS

Hydrocarbons are organic molecules consisting entirely of carbon and hydrogen, such as methane (CH_4) described above. We often use hydrocarbons in our daily lives as fuels—like the propane in a gas grill or the butane in a lighter. The many covalent bonds between the atoms in hydrocarbons store a great amount of energy, which is released when these molecules are burned (oxidized). Methane, an excellent fuel, is the simplest hydrocarbon molecule, with a central carbon atom bonded to four different hydrogen atoms

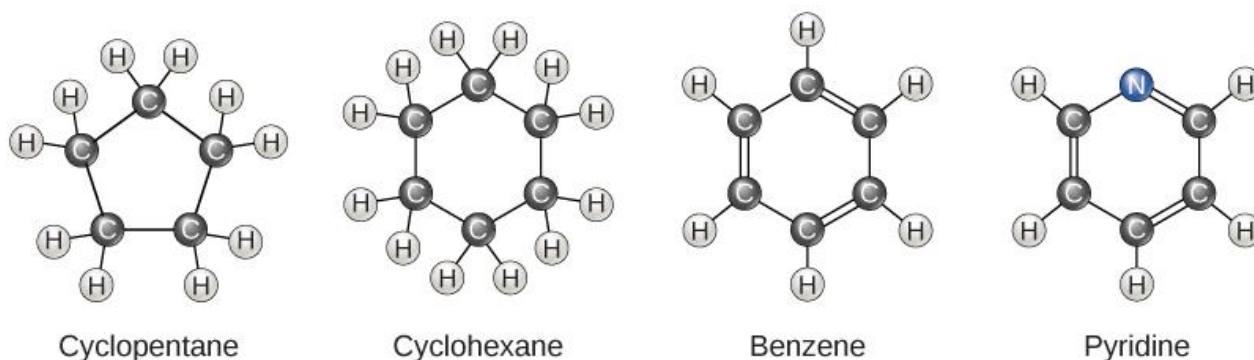
HYDROCARBON CHAINS

Hydrocarbon chains are formed by successive bonds between carbon atoms and may be branched or unbranched. Furthermore, the overall geometry of the molecule is altered by the different geometries of single, double, and triple covalent bonds. The hydrocarbons ethane, ethene, and ethyne serve as examples of how different carbon-to-carbon bonds affect the geometry of the molecule. The names of all three molecules start with the prefix “eth-,” which is the prefix for two carbon hydrocarbons. The suffixes “-ane,” “-ene,” and “-yne” refer to the presence of single, double, or triple carbon-carbon bonds, respectively. Thus, propane, propene, and propyne follow the same pattern with three carbon molecules, butane, butene, and butyne for four carbon molecules, and so on. Double and triple bonds change the geometry of the molecule: single bonds allow rotation along the axis of the bond, whereas double bonds lead to a planar configuration and triple bonds to a linear on

Methane (CH ₄)	Ethane (C ₂ H ₆)	Ethene (C ₂ H ₄)
		
Tetrahedral (single bond)	Tetrahedral (single bond)	Planar (double bond)

HYDROCARBON RINGS

So far, the hydrocarbons we have discussed have been aliphatic hydrocarbons, which consist of linear chains of carbon atoms. Another type of hydrocarbon, aromatic hydrocarbons, consists of closed rings of carbon atoms. Ring structures are found in hydrocarbons, sometimes with the presence of double bonds, which can be seen by comparing the structure of cyclohexane to benzene



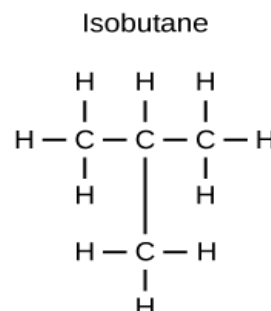
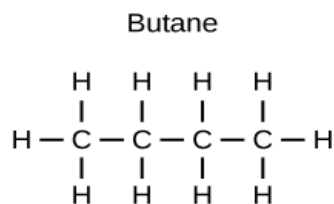
Carbon can form five- and six-membered rings. Single or double bonds may connect the carbons in the ring, and nitrogen may be substituted for carbon.

ISOMERS

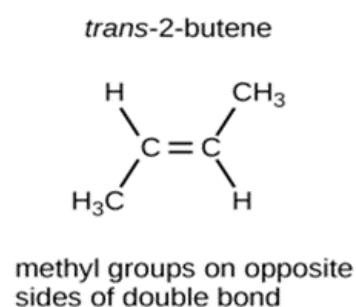
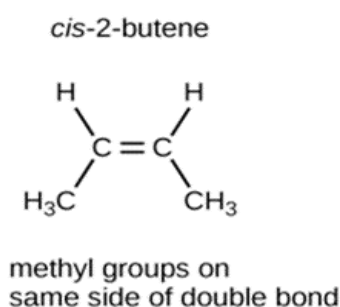
The three-dimensional placement of atoms and chemical bonds within organic molecules is central to understanding their chemistry. Molecules that share the same chemical formula but differ in the placement (structure) of their atoms and/or chemical bonds are known as isomers. Structural isomers (like butane and isobutene shown in below) differ in the placement of their covalent bonds: both molecules have four carbons and ten hydrogens (C₄H₁₀), but the different arrangement of the atoms within the molecules leads to differences in their chemical properties. For example, due to their different chemical properties, butane is suited for use as a

fuel for cigarette lighters and torches, whereas isobutene is suited for use as a refrigerant and a propellant in spray cans. Geometric isomers, on the other hand, have similar placements of their covalent bonds but differ in how these bonds are made to the surrounding atoms, especially in carbon-to-carbon double bonds. In the simple molecule butene (C_4H_8), the two methyl groups (CH_3) can be on either side of the double covalent bond central to the molecule, as illustrated below

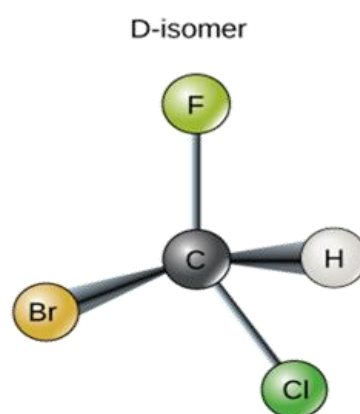
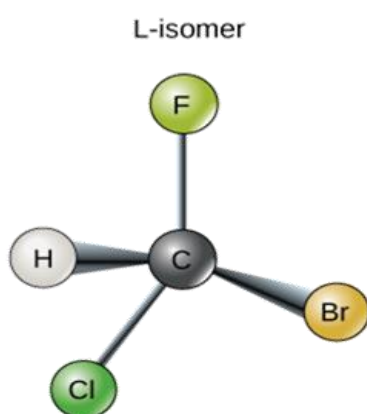
(a) Structural isomers



(b) Geometric isomers



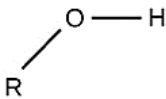
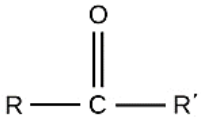
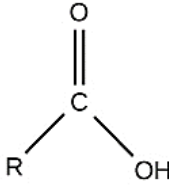
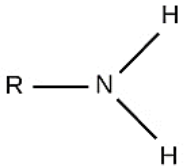
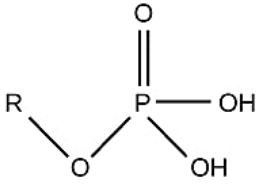
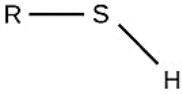
(c) Enantiomers

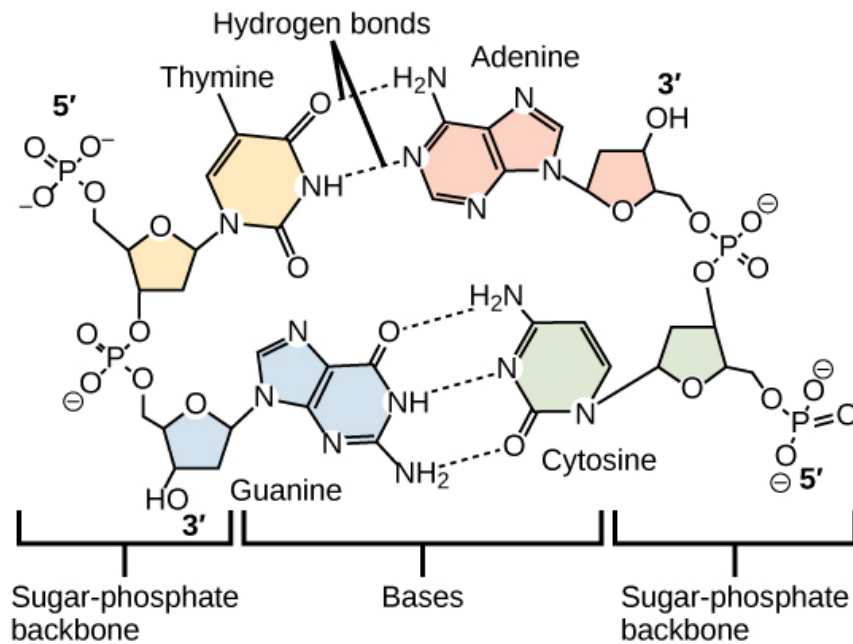


FUNCTIONAL GROUPS

Functional groups are groups of atoms that occur within molecules and confer specific chemical properties to those molecules. They are found along the “carbon backbone” of macromolecules. This carbon backbone is formed by chains and/or rings of carbon atoms with

the occasional substitution of an element such as nitrogen or oxygen. Molecules with other elements in their carbon backbone are substituted hydrocarbons. The functional groups in a macromolecule are usually attached to the carbon backbone at one or several different places along its chain and/or ring structure. Each of the four types of macromolecules—proteins, lipids, carbohydrates, and nucleic acids—has its own characteristic set of functional groups that contributes greatly to its differing chemical properties and its function in living organisms. A functional group can participate in specific chemical reactions they include: hydroxyl, methyl, carbonyl, carboxyl, amino, phosphate, and sulfhydryl. These groups play an important role in the formation of molecules like DNA, proteins, carbohydrates, and lipids. Functional groups are usually classified as hydrophobic or hydrophilic depending on their charge or polarity characteristics.

Functional Group	Structure	Properties
Hydroxyl		Polar
Methyl	$R - CH_3$	Nonpolar
Carbonyl		Polar
Carboxyl		Charged, ionizes to release H^+ . Since carboxyl groups can release H^+ ions into solution, they are considered acidic.
Amino		Charged, accepts H^+ to form NH_3^+ . Since amino groups can remove H^+ from solution, they are considered basic.
Phosphate		Charged, ionizes to release H^+ . Since phosphate groups can release H^+ ions into solution, they are considered acidic.
Sulfhydryl		Polar

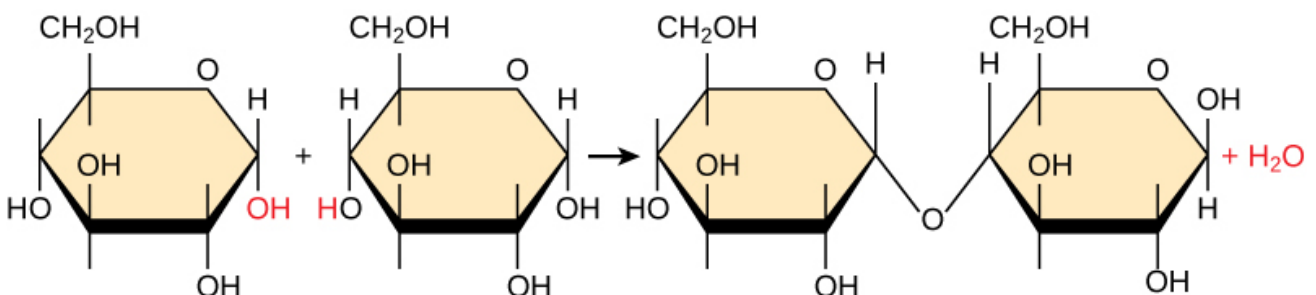


Hydrogen bonds connect two strands of DNA together to create the double-helix structure

SYNTHESIS OF BIOLOGICAL MACROMOLECULES

Dehydration Synthesis

Most macromolecules are made from single subunits, or building blocks, called monomers. The monomers combine with each other using covalent bonds to form larger molecules known as polymers. In doing so, monomers release water molecules as byproducts. This type of reaction is known as dehydration synthesis, which means “to put together while losing water.”

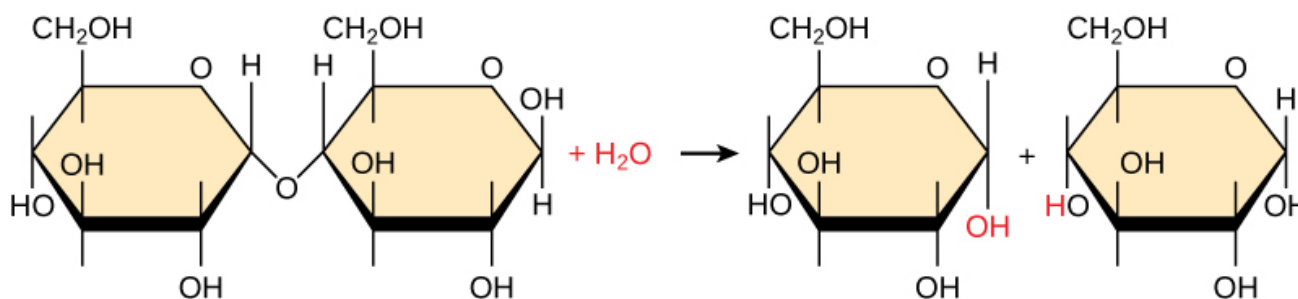


In the dehydration synthesis reaction depicted above, two molecules of glucose are linked together to form the disaccharide maltose. In the process, a water molecule is formed. In a dehydration synthesis reaction, the hydrogen of one monomer combines with the hydroxyl group of another monomer, releasing a molecule of water. At the same time, the monomers share electrons and form covalent bonds. As additional monomers join, this chain of repeating

monomers forms a polymer. Different types of monomers can combine in many configurations, giving rise to a diverse group of macromolecules. Even one kind of monomer can combine in a variety of ways to form several different polymers: for example, glucose monomers are the constituents of starch, glycogen, and cellulose.

Hydrolysis

Polymers are broken down into monomers in a process known as hydrolysis, which means “to split water,” a reaction in which a water molecule is used during the breakdown. During these reactions, the polymer is broken into two components: one part gains a hydrogen atom (H^+) and the other gains a hydroxyl molecule (OH^-) from a split water molecule.



In the hydrolysis reaction shown here, the disaccharide maltose is broken down to form two glucose monomers with the addition of a water molecule. Note that this reaction is the reverse of the synthesis reaction. Dehydration and hydrolysis reactions are catalyzed, or “sped up,” by specific enzymes; dehydration reactions involve the formation of new bonds, requiring energy, while hydrolysis reactions break bonds and release energy. These reactions are similar for most macromolecules, but each monomer and polymer reaction is specific for its class. For example, in our bodies, food is hydrolyzed, or broken down, into smaller molecules by catalytic enzymes in the digestive system. This allows for easy absorption of nutrients by cells in the intestine. Each macromolecule is broken down by a specific enzyme. For instance, carbohydrates are broken down by amylase, sucrase, lactase, or maltase. Proteins are broken down by the enzymes pepsin and peptidase, and by hydrochloric acid. Lipids are broken down by lipases. Breakdown of these macromolecules provides energy for cellular activities.