Part 1: Introduction to Biochemistry (Pages 1–17)

This part of your lecture introduces **what biochemistry is**, **what makes something alive**, the **characteristics of living systems**, and the **basic structure and function of cells**. It also includes several clinical scenarios to help you apply these concepts.

Page 1 – Introduction to Biochemistry

The title of the lecture is **"Introduction to Biochemistry."**

What is Biochemistry?

Biochemistry is the branch of science that studies the **chemical substances** found in living organisms and the **chemical reactions** that allow life to exist.

Think about your body.

Everything inside you is made of chemicals:

- Water
- Proteins
- DNA
- Carbohydrates
- Lipids (fats)
- Minerals
- Vitamins

Biochemistry explains:

- how muscles contract,
- how food is converted into energy,
- how wounds heal,
- how medicines work,
- and why diseases occur.

Why is biochemistry important in nursing?

As a nurse, almost everything you do is related to biochemistry.

For example:

A diabetic patient receives insulin.

Biochemistry explains:

- how insulin lowers blood glucose,
- how cells take in glucose,
- and what happens if insulin is absent.

Without biochemistry, it would be difficult to understand why treatments and medications work.

Pages 2–3 – "Life Ceases Without Molecules"

The lecture asks an interesting question:

"If I gave you a jar containing every single molecule that makes up your body (every atom, every protein, every strand of DNA), but they were just sitting there, unorganized, not interacting, would that jar be alive?"

The answer is:

No.

Why not?

Many people think life is simply made of molecules.

That is true.

But molecules alone are **not enough**.

Imagine taking apart a cellphone.

Inside the phone are:

- battery
- wires
- chips
- camera
- screen

Now place every piece into a box.

Does the phone still work?

No.

Why?

Because the parts are no longer connected.

Exactly the same thing happens in living organisms.

Your body contains:

- DNA
- proteins
- water
- fats
- minerals
- enzymes

If these molecules are simply lying in a jar, nothing happens.

There is:

- no heartbeat,
- no breathing,
- no metabolism,
- no cell division,
- no thinking.

The lecture summarizes this idea by stating that **molecules are necessary but not sufficient** for life.

What actually makes something alive?

Living organisms require **organization**.

Cells are highly organized.

Inside every cell:

DNA gives instructions.

↓

Ribosomes make proteins.

↓

Proteins become enzymes.

↓

Enzymes perform chemical reactions.

↓

Energy is produced.

↓

The cell survives.

All of these events occur in a coordinated manner.

Life depends not only on the presence of molecules but also on how those molecules interact.

Page 4 – Four Key Themes of Life

The lecture introduces four major themes:

- Hierarchical organization
- Energy flow
- Stereospecificity
- Self-regulation

Let's discuss each one.

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1. Hierarchical Organization

"Hierarchy" means **levels**.

Living organisms are built from smaller structures that combine to form larger ones.

For example:

Atoms

↓

Molecules

↓

Macromolecules

↓

Organelles

↓

Cells

↓

Tissues

↓

Organs

↓

Organ systems

↓

Human body

Each level depends on the level below it.

Example

A heart is made of heart muscle tissue.

Heart muscle tissue is made of cardiac muscle cells.

Cells contain organelles.

Organelles contain molecules.

Without molecules, there would be no organelles.

Without organelles, there would be no cells.

Without cells, there would be no heart.

---

2. Energy Flow

The lecture asks:

"If I took away only the energy flow (cut off a cell's energy supply completely, but left everything else intact), what happens and why?"

The next slide answers:

Molecular order breaks down; the cell cannot maintain its organization or repair itself, so it dies because maintaining a highly ordered structure requires constant energy.

Why?

Every activity inside a cell requires energy.

Examples:

- pumping sodium and potassium,
- making proteins,
- repairing DNA,
- muscle contraction,
- nerve impulses.

The energy currency of the cell is **ATP (adenosine triphosphate)**.

Without ATP:

- proteins stop functioning,
- pumps stop working,
- waste accumulates,
- the cell loses organization,
- eventually the cell dies.

Think of ATP like electricity in a hospital.

The equipment is still present, but without power it cannot function.

---

3. Is a Dead Body Still Organized?

The lecture asks:

"Is a dead body still organized the same way a living one is?"

The answer is:

Not for long.

Immediately after death, the body still looks organized.

However:

- ATP production stops.
- Cells cannot repair damage.
- Membranes begin to fail.
- Enzymes start breaking down the body's own tissues (autolysis).
- Bacteria begin decomposition.

The organized state gradually disappears because the body can no longer maintain itself.

---

4. Can Chemistry Explain Life?

The lecture asks:

"A classmate says, 'Life has some mysterious energy that chemistry can't explain.' How would you respond?"

Based on the lecture, life does **not** require a mysterious force.

Instead, life is explained by:

- organized molecules,
- continuous chemical reactions,
- energy flow,
- regulation,
- and interactions between biomolecules.

Living cells follow the same chemical laws as non-living matter, but they are organized in a highly coordinated way.

Pages 8–17 – Cell Organelles

The lecture now shifts to the cell and asks you to diagnose which organelle is malfunctioning based on different patient scenarios.

Scenario 1 – Golgi Apparatus

The patient:

Proteins are made, but insulin is never packaged and exported from pancreatic cells.

What happened?

The **Golgi apparatus** is malfunctioning.

What does the Golgi do?

Think of the Golgi as a **shipping center**.

Proteins are produced in the rough ER.

They are then sent to the Golgi.

The Golgi:

- modifies proteins,
- sorts them,
- packages them,
- sends them to their final destination.

If the Golgi fails:

Proteins are still made, but they cannot be delivered.

Result

Insulin remains trapped inside the cell.

Without insulin secretion:

- blood glucose rises,
- cells cannot absorb glucose efficiently,
- symptoms similar to diabetes may occur.

Scenario 2 – Lysosome

The patient cannot break down worn-out cellular components, and waste accumulates inside cells.

The damaged organelle is the **lysosome**.

Lysosomes contain digestive enzymes that:

- destroy bacteria,
- digest damaged organelles,
- recycle cellular materials.

Think of lysosomes as the **cell's recycling and garbage disposal system**.

If lysosomes stop working:

- waste accumulates,

- damaged organelles remain inside the cell,
- the cell gradually becomes dysfunctional.

Scenario 3 – Smooth Endoplasmic Reticulum (SER)

A patient's liver cells cannot detoxify a new medication.

The affected organelle is the **smooth endoplasmic reticulum**.

The SER:

- synthesizes lipids,
- stores calcium,
- detoxifies drugs and harmful chemicals, especially in liver cells.

If the SER is impaired:

- medications remain in the body longer,
- toxins accumulate,
- adverse drug effects become more likely.

Scenario 4 – Mitochondria

A muscle cell has enough glucose and oxygen but still produces very little ATP.

The problem is the **mitochondria**.

Mitochondria are the **powerhouses of the cell**.

They convert energy from nutrients into ATP through cellular respiration.

If mitochondria malfunction:

- ATP production decreases,
- muscles become weak,
- fatigue develops because the cell lacks usable energy.

Scenario 5 – Antibiotics

The lecture asks why antibiotics kill bacteria without damaging human cells.

This is because bacteria are **prokaryotes**, while human cells are **eukaryotes**.

Many antibiotics target structures found only in bacteria, such as:

- the bacterial cell wall,
- or the bacterial 70S ribosome.

Human cells lack these bacterial structures, so the antibiotics can destroy bacteria while leaving human cells largely unaffected. This principle is called **selective toxicity**.

Page 15 – Stereospecificity

The lecture compares stereospecificity to a **left glove fitting only a left hand**.

Enzymes and receptors recognize molecules based on **shape**.

If the shape does not match:

- the enzyme cannot bind,
- or the drug may not work correctly.

This is why some drugs have one beneficial form and another that is ineffective or harmful.

Page 16 – Why Cells Have Organelles

The lecture asks why reactions are separated into different organelles instead of happening in one large space.

The answer is **compartmentalization**.

Each organelle creates a specialized environment for specific reactions.

For example:

- Lysosomes contain acidic digestive enzymes.
- Mitochondria specialize in ATP production.
- The Golgi modifies and packages proteins.

If all reactions occurred in the same place, they could interfere with one another, reducing the cell's efficiency and damaging important molecules.

Part 1 Summary

By the end of this section, you should understand that:

- Life depends on **organized, interacting molecules**, not just the molecules themselves.
- Cells require a constant supply of **energy (ATP)** to maintain order and survive.
- Organelles act like specialized departments within a factory, each with a unique function.
- Different organelles produce different symptoms when they malfunction.
- Antibiotics work because they target structures unique to bacteria.
- The **shape** of molecules is crucial for biological recognition.
- Compartmentalization allows many chemical reactions to occur efficiently at the same time without interfering with one another.

PART 2 – WATER (Pages 18–28)

This section explains **why water is the most important molecule in the human body**. Nearly **60–70% of the adult human body is water**, and almost every biochemical reaction occurs in a watery environment. The lecture focuses on **the structure of water** and shows how that structure explains nearly all of water's unique properties.

Page 18 – Why is Water's Bond Angle 104.5° Instead of 109.5°?

The lecture asks:

"Why is water's angle 104.5° instead of 109.5°, and why does that matter?"

Let's answer this step by step.

Step 1: What is Water Made Of?

A water molecule has:

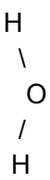
- **1 oxygen atom (O)**
- **2 hydrogen atoms (H)**

Chemical formula:

H_2O

The oxygen atom is in the middle, while the two hydrogen atoms are attached to it.

You can imagine it like this:



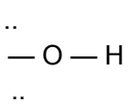
However, water is **not a straight line**.

Instead, it has a **bent (angular)** shape.

Step 2: Oxygen Has Four Regions of Electrons

Around the oxygen atom there are four regions of electron density:

- Two bonding pairs (shared with hydrogen atoms)
- Two lone pairs (not shared)



The two dots (. .) represent the lone pairs.

Step 3: Why Isn't the Angle 109.5°?

According to the **VSEPR (Valence Shell Electron Pair Repulsion) Theory**, electron pairs repel each other.

Imagine four people standing in a small room.

Everyone wants as much personal space as possible.

They spread out evenly.

With four electron regions, the ideal arrangement is **109.5°**, which is called a **tetrahedral arrangement**.

If all four regions were bonding pairs, the angle would remain 109.5°.

Step 4: Lone Pairs Push Harder

Lone pairs occupy more space than bonding pairs because they belong entirely to the oxygen atom.

As a result, lone pairs repel the bonding pairs **more strongly**.

Think of it like this:

Imagine four balloons tied together.

Two balloons are larger (the lone pairs).

The larger balloons push the smaller balloons closer together.

The hydrogen atoms are pushed inward.

Instead of:

109.5°

the angle becomes:

104.5°

That is the actual bond angle of water.

Why Does This Matter?

Many students memorize **104.5°** without understanding why it is important.

The smaller angle creates the bent shape.

The bent shape causes an uneven distribution of charge.

That uneven distribution makes water **polar**.

If water were perfectly linear, many of its special biological properties would disappear.

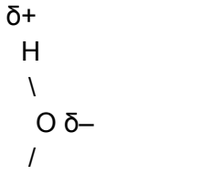
Water is a Polar Molecule

Oxygen attracts electrons much more strongly than hydrogen.

This property is called **electronegativity**.

Because oxygen is more electronegative:

- electrons spend more time near oxygen,
- oxygen becomes slightly negative (δ−),
- each hydrogen becomes slightly positive (δ+).



H
δ+

Water now has:

- one positive side,
- one negative side.

This is called a **dipole**.

What Does "Polar" Mean?

A polar molecule has:

- one end that is slightly positive,
- another end that is slightly negative.

Think of water as a tiny magnet.

Just as a magnet has:

North

South

Water has:

Positive

Negative

These opposite charges attract other molecules.

This single property—**polarity**—explains almost everything about water's behavior.

Hydrogen Bonding

Because water molecules are polar, they attract each other.

The positive hydrogen of one water molecule is attracted to the negative oxygen of another.

Hδ+ — Oδ− ····· Hδ+ — Oδ−
↑
Hydrogen bond

Notice that the hydrogen bond is **not inside the molecule**.

It is **between two different water molecules**.

Covalent Bond vs Hydrogen Bond

Many students confuse these two.

Covalent Bond

Occurs **inside one water molecule**.

Oxygen shares electrons with hydrogen.

Very strong.

H — O — H

Hydrogen Bond

Occurs **between neighboring water molecules**.

No electrons are shared.

It is an attraction caused by partial charges.

Much weaker than a covalent bond.

H_2O ····· H_2O

Types of Intermolecular Forces

The handout explains three intermolecular forces.

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1. London Dispersion Forces

These occur because electrons are always moving.

As electrons move, they temporarily create tiny positive and negative regions.

Even nonpolar molecules experience this temporary attraction.

Examples:

- Helium
- Oxygen gas
- Carbon dioxide

These are the **weakest** intermolecular forces.

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2. Dipole–Dipole Interactions

These occur between **polar molecules**.

The positive end of one molecule attracts the negative end of another.

Example:

Hydrogen chloride (HCl)

or

Water molecules.

These interactions are stronger than London dispersion forces.

3. Hydrogen Bonding

Hydrogen bonding is a particularly strong type of dipole–dipole interaction.

It occurs only when hydrogen is bonded to:

- Oxygen (O)
- Nitrogen (N)
- Fluorine (F)

Remember the shortcut:

HONF

Hydrogen bonded to **O, N, or F** can form hydrogen bonds.

Among the three intermolecular forces discussed in this chapter, hydrogen bonding is the **strongest**.

One Water Molecule Can Form Four Hydrogen Bonds

Each water molecule has:

- two hydrogen atoms that can **donate** hydrogen bonds,
- two lone pairs on oxygen that can **accept** hydrogen bonds.

Total:

4 hydrogen bonds

Imagine one person holding hands with four other people.

This creates a giant network.

Liquid water behaves like a constantly changing network of connected molecules.

Why Does Water Have a High Boiling Point?

Normally, small molecules boil at relatively low temperatures.

Water is different.

Before water can become steam:

Hydrogen bonds between molecules must first be broken.

Breaking these bonds requires extra heat.

That is why water boils at **100°C**, much higher than expected for such a small molecule. This property also helps explain evaporative cooling, such as sweating.

Why Does Sweating Cool the Body?

Sweat is mostly water.

The fastest-moving water molecules escape first during evaporation.

These molecules carry away heat energy.

The remaining water—and your skin—becomes cooler.

This is why sweating lowers body temperature.

Why Does Ice Float?

The lecture asks:

"Why does ice float in a water pitcher?"

Most solids are denser than their liquids.

Water is unusual.

When water freezes:

Hydrogen bonds become fixed in a rigid **hexagonal lattice**.

The molecules are held farther apart than they are in liquid water.

Because the molecules are farther apart:

The same mass occupies a larger volume.

This lowers the density.

Since ice is less dense than liquid water, it floats.

This property is important because lakes freeze from the top down, insulating the water below and allowing aquatic life to survive.

Water Dissolves Salt

The lecture asks:

"Using the idea of an ion–dipole interaction, explain what happens when NaCl dissolves into a patient's IV fluid."

Table salt is:

NaCl

When placed in water:

NaCl separates into:

- Na^+ (sodium ion)
- Cl^- (chloride ion)

Water molecules surround each ion.

Around Na^+:

The oxygen end (δ−) points toward the positive sodium ion.

Around Cl^-:

The hydrogen ends (δ+) point toward the negative chloride ion.

This attraction is called an **ion–dipole interaction**.

When enough water molecules surround the ions, they pull them away from the crystal and keep them separated in solution.

This is why saline IV fluids contain dissolved sodium and chloride ions that can move freely through the bloodstream.

Water Dissolves Polar Molecules

Water can also dissolve many **polar covalent molecules**, such as alcohols and sugars.

Unlike salt, these molecules do not separate into ions.

Instead, water forms **hydrogen bonds** directly with them.

This is why substances like glucose travel dissolved in blood plasma.

Water is the Universal Solvent

The lecture explains that water's ability to hydrate ions, polar molecules, and large biomolecules makes it the body's **universal solvent and transport medium**.

Water transports:

- oxygen (mostly indirectly through blood components),
- nutrients,
- electrolytes,
- hormones,
- waste products,
- medicines.

Without water, these substances could not move efficiently throughout the body.

Page 27 – The Big Idea

The lecture asks:

"All three questions trace back to one property of water. What is it?"

The answer is:

Water is polar.

Because water is polar:

- it forms hydrogen bonds,
- it has a high boiling point,
- ice is less dense than liquid water,
- it dissolves ions,
- it dissolves many polar molecules,
- and it serves as the body's universal solvent.

Almost every important biological property of water can be traced back to its **polarity**.

⭐ Part 2 Summary (Must Know for Exams)

- Water has a **bent shape** because lone pairs repel bonding pairs, reducing the bond angle from **109.5° to 104.5°**.
- Water is **polar** because oxygen is more electronegative than hydrogen.
- Water molecules attract one another through **hydrogen bonds**.
- The three intermolecular forces discussed are **London dispersion < dipole–dipole < hydrogen bonding** (weakest to strongest).
- One water molecule can form **up to four hydrogen bonds**.
- Hydrogen bonding explains water's **high boiling point** and why **sweating cools the body**.
- Ice floats because hydrogen bonds create an open structure that makes ice **less dense** than liquid water.
- Water dissolves ionic compounds through **ion–dipole interactions** and dissolves many polar molecules through **hydrogen bonding**.
- **Water's polarity** is the key property that explains almost all of its biological importance.

PART 3 – BONDING PROPERTIES OF CARBON (Pages 29–36)

This section explains **why carbon is called the backbone of life**. Although carbon makes up **less than 1% of the Earth's crust**, every major biological molecule—carbohydrates, proteins, lipids, nucleic acids, hormones, and vitamins—is built around carbon. The lecture explains that this is because of **how carbon bonds**, not because of how abundant it is.

Page 29 – Why is Carbon the Backbone of Biology?

The lecture asks:

"Carbon isn't even 1% of Earth's crust, yet it's the backbone of biology. What is it specifically about carbon's bonding that earns it that role?"

At first, this seems strange.

If carbon is not very abundant, why is it used to build every living organism?

The answer is **its bonding properties**.

What Makes Carbon Special?

Carbon has **4 valence electrons**.

What are valence electrons?

Valence electrons are the electrons in the **outermost shell** of an atom.

These are the electrons that participate in chemical bonding.

Carbon has:

- Atomic number = 6
- Electron arrangement = **2,4**

The first shell holds 2 electrons.

The second (outer) shell has **4 electrons**.

Those 4 electrons are its **valence electrons**.

Why Does Carbon Form Four Bonds?

Atoms are generally more stable when their outer shell contains **8 electrons** (the octet rule).

Carbon has only **4**.

So it needs **4 more electrons**.

Instead of gaining or losing four electrons—which would require a lot of energy—carbon **shares electrons** with other atoms.

Sharing electrons creates **covalent bonds**.

Because carbon needs four more electrons, it usually forms **four covalent bonds**.

The lecture calls this property **tetravalency**.

What is Tetravalency?

"Tetra" means **four**.

"Valency" refers to the number of bonds an atom forms.

So **tetravalency** means:

Carbon normally forms **four covalent bonds**.

This is one of the most important ideas in organic chemistry.

Carbon Can Bond to Many Different Atoms

Carbon doesn't only bond with hydrogen.

It can also bond with:

- Carbon (C)
- Oxygen (O)
- Nitrogen (N)
- Sulfur (S)
- Phosphorus (P)

Because of this, carbon can produce millions of different compounds.

Carbon Can Bond to Itself

This is perhaps carbon's most remarkable property.

Imagine if people could only hold hands with one type of person.

Building large groups would be difficult.

Carbon is different.

A carbon atom can bond directly to another carbon atom.

Then another.

Then another.

This allows carbon to form:

Straight chains

C–C–C–C–C

Branched chains

```
    C
    |
C–C–C–C
```

Rings

```
  C──C
 /    \
C      C
 \    /
  C──C
```

These different arrangements create an enormous variety of organic molecules.

Why Can Living Organisms Build So Many Molecules?

Because carbon can:

- form four bonds,
- bond repeatedly to itself,
- create chains,
- create branches,
- create rings,
- bond with many other elements.

That is why millions of organic compounds exist.

Examples include:

- glucose,
- amino acids,
- proteins,
- DNA,
- cholesterol,
- hormones,
- fats.

All of these are carbon-based.

Pages 30–33 – Carbon Bonding and Molecular Shape

The lecture shows that carbon can form different combinations of single, double, and triple bonds, and that these combinations determine the molecule's **geometry**.

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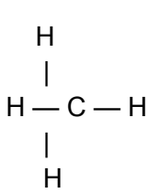
1. Four Single Bonds

Example:

Methane (CH_4)



Carbon forms:

- four single bonds.

This arrangement creates a **tetrahedral** geometry.

Bond angle:

109.5°

Imagine a three-dimensional pyramid where the bonds spread out as far apart as possible.

Why 109.5°?

Each bonding pair repels the others.

To reduce this repulsion, the bonds spread out evenly.

The ideal angle is:

109.5°

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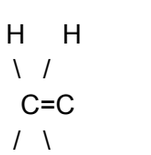
2. One Double Bond + Two Single Bonds

Example:

Ethene (C_2H_4)



H H

Now carbon has:

- one double bond,
- two single bonds.

The shape changes.

Instead of tetrahedral, it becomes:

Trigonal planar

Bond angle:

120°

Everything lies in one flat plane.

---

3. One Triple Bond

Example:

Ethyne (C_2H_2)

H–C≡C–H

Carbon forms:

- one triple bond,
- one single bond.

This creates a **linear** geometry.

Bond angle:

180°

The atoms form a straight line.

Easy Way to Remember

Bonding around Carbon	Shape	Bond Angle
4 single bonds	Tetrahedral	**109.5°**
1 double + 2 single bonds	Trigonal planar	**120°**
1 triple bond	Linear	**180°**

These three bond angles are commonly tested in chemistry and biochemistry.

Why Do These Shapes Matter?

The three-dimensional shape of a molecule determines:

- how it interacts with other molecules,
- how enzymes recognize it,
- how receptors bind it,
- how medicines work.

Two molecules with the same atoms but different shapes may behave very differently in the body.

Page 34 – Hydrocarbons

The lecture defines a **hydrocarbon** as the simplest type of organic molecule, containing **only carbon and hydrogen atoms**.

Hydrocarbons form the basic framework of many larger biological molecules.

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1. Saturated Hydrocarbons (Alkanes)

The lecture states that saturated hydrocarbons contain **only single bonds** between carbon atoms.

Example:

Propane

$CH_3–CH_2–CH_3$

Because all carbon-carbon bonds are single bonds, each carbon has the maximum possible number of hydrogen atoms.

These molecules are called **saturated**.

---

2. Unsaturated Hydrocarbons

Unsaturated hydrocarbons contain:

- one or more **double bonds** (alkenes),
- or one or more **triple bonds** (alkynes).

Examples:

Ethene

$CH_2=CH_2$

Ethyne

$HC\equiv CH$

Since double and triple bonds reduce the number of hydrogen atoms attached, these molecules are called **unsaturated**.

Why Are Unsaturated Molecules Important?

Unsaturated bonds affect:

- the shape of molecules,
- their flexibility,
- and their chemical reactivity.

This becomes especially important later when studying lipids and fatty acids.

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3. Aromatic Hydrocarbons

The lecture also mentions **aromatic hydrocarbons**, such as **benzene**, which contain ring systems with alternating or delocalized double bonds.

Benzene is represented as a six-carbon ring.

Its electrons are shared around the ring, making it unusually stable compared with many other unsaturated compounds.

Page 35 – Classification of Carbon Atoms

The lecture introduces the idea that each carbon atom in a chain can be classified based on **how many other carbon atoms it is directly attached to**.

For example:

Primary (1°) Carbon

Attached to **one** other carbon.

Secondary (2°) Carbon

Attached to **two** other carbons.

Tertiary (3°) Carbon

Attached to **three** other carbons.

Quaternary (4°) Carbon

Attached to **four** other carbons.

This system helps chemists describe the structure of organic molecules.

Page 36 – Functional Groups (Introduction)

The final page in this section introduces **functional groups**.

The lecture defines a functional group as:

A specific arrangement of atoms attached to the carbon skeleton that determines how the molecule behaves and reacts.

Think of the carbon skeleton as the "body" of the molecule.

The functional group is like a special attachment that gives the molecule its personality.

For example:

- An alcohol contains a **hydroxyl (-OH)** group.

- A carboxylic acid contains a **carboxyl (-COOH)** group.
- An amine contains an **amino (-NH_2)** group.

Even if two molecules have similar carbon skeletons, different functional groups can make one acidic, another basic, another soluble in water, or another chemically reactive.

The next section of your lecture focuses entirely on these functional groups.

⭐ Part 3 Summary (Must Know for Exams)

- Carbon is the backbone of life because of its **tetravalency**—it has **4 valence electrons** and normally forms **4 covalent bonds**.
- Carbon can bond to itself repeatedly, allowing it to form **chains, branches, and rings**.
- Carbon also bonds with elements such as **H, O, N, S, and P**, creating an enormous variety of biological molecules.
- Carbon's bonding determines molecular geometry:
  - **4 single bonds → tetrahedral (109.5°)**
  - **1 double + 2 single bonds → trigonal planar (120°)**
  - **1 triple bond → linear (180°)**
- Hydrocarbons contain only **carbon and hydrogen** and are classified as **saturated (single bonds)**, **unsaturated (double or triple bonds)**, or **aromatic (ring structures like benzene)**.
- A **functional group** is the part of an organic molecule that determines its chemical properties and reactions, making it one of the most important concepts in organic chemistry.

PART 4 – FUNCTIONAL GROUPS (Pages 37–54)

This part of your lecture introduces **functional groups**, one of the most important topics in biochemistry. The lecture defines a functional group as **a specific arrangement of atoms attached to the carbon skeleton that determines how the molecule behaves and reacts.**

Think of it this way:

Imagine you have two identical cars.

One has police lights.

The other has an ambulance siren.

Although both cars have the same basic body, they perform different jobs because of what is attached to them.

Organic molecules are similar.

The **carbon skeleton** is like the body of the car.

The **functional group** determines the molecule's properties and chemical behavior.

Why are Functional Groups Important?

Without functional groups, most organic molecules would behave almost the same way.

Adding a functional group changes a molecule's:

- chemical reactivity,
- acidity or basicity,
- solubility in water,
- smell,
- boiling point,
- biological function.

This is why two molecules with almost identical carbon skeletons can have completely different functions in the body.

HYDROXYL GROUP (–OH)

The lecture introduces the hydroxyl group as one of the common oxygen-containing functional groups.

Structure

R–OH

R = carbon chain

OH = hydroxyl group

What is it?

A hydroxyl group contains:

- one oxygen atom
- one hydrogen atom

bonded together.

Do **not** confuse this with hydroxide (OH^-).

The hydroxyl group is **part of an organic molecule**, while hydroxide is an ion.

What happens when a hydroxyl group is added?

Because oxygen is very electronegative, the hydroxyl group is **polar**.

This allows the molecule to form hydrogen bonds with water.

As a result:

- water solubility increases,
- boiling point increases,
- the molecule becomes more hydrophilic ("water-loving").

Biological Examples

- Ethanol
- Glycerol
- Sugars like glucose (which contain several hydroxyl groups)

Because glucose has many –OH groups, it dissolves easily in blood.

Nursing Connection

When a patient receives IV glucose, the glucose remains dissolved in blood because of its many hydroxyl groups.

ETHER GROUP

The lecture also includes the ether functional group among the oxygen-containing groups.

Structure

R–O–R'

One oxygen atom connects two carbon-containing groups.

Properties

Unlike alcohols:

- ethers cannot donate hydrogen bonds,
- but oxygen can accept hydrogen bonds.

Therefore,

ethers are usually **less soluble in water** than alcohols.

Example

Diethyl ether

Historically, ether was used as a surgical anesthetic.

ALDEHYDE

The lecture next discusses carbonyl-containing functional groups.

What is a Carbonyl?

A carbonyl group is:

C=O

One carbon atom

double-bonded

to one oxygen atom.

Aldehyde Structure

R–CHO

The carbonyl carbon is located at the **end** of the carbon chain.

Example

Glucose contains an aldehyde group in its open-chain form.

Why is it Important?

Aldehydes are reactive.

They participate in many metabolic reactions.

KETONE

Ketones also contain a carbonyl group.

However,

the carbonyl carbon is located **inside** the carbon chain.

Structure

R–CO–R'

Notice that carbon is attached to carbon atoms on **both sides**.

Example

Acetone

Ketone bodies (important in diabetes and fasting)

Clinical Importance

Patients with uncontrolled diabetes may produce excessive ketone bodies.

These can lead to:

- ketoacidosis,
- dehydration,
- rapid breathing.

CARBONYL GROUP

The lecture specifically defines the carbonyl group as **an oxygen atom double-bonded to a carbon atom.**

Remember:

Carbonyl is **not** a complete functional group by itself in this context.

Instead,

it forms part of:

- aldehydes,
- ketones,
- carboxylic acids,
- esters,
- amides.

CARBOXYL GROUP

The lecture states that the carboxyl group **consists of a carbonyl group (C=O) with a hydroxyl group (–OH) bonded to the carbonyl carbon atom.**

Structure

–COOH

It contains:

- one carbonyl group,
- one hydroxyl group.

Why is it Special?

Carboxyl groups can release hydrogen ions (H^+).

That makes them **acidic**.

Biological Examples

- Fatty acids
- Amino acids
- Citric acid
- Lactic acid

Nursing Importance

Lactic acid accumulates during poor oxygen supply.

Too much lactic acid may lead to **lactic acidosis**, lowering blood pH.

CARBOXYLIC ACID

The lecture defines a carboxylic acid as **an organic compound containing a carboxyl group (–COOH)** and gives the general formula **R–COOH**.

Examples include:

- Acetic acid (vinegar)
- Formic acid
- Fatty acids

These compounds are acidic because they can donate H^+ ions.

ESTER

The lecture includes esters among the oxygen-containing functional groups.

Structure

R–COO–R'

Esters form when:

Carboxylic acid

-

Alcohol

↓

Ester

-

Water

This reaction is called **esterification**.

Biological Importance

Esters are found in:

- fats (triglycerides),
- phospholipids,
- many natural fragrances.

Many fruits smell sweet because of esters.

AMINE

The lecture defines amines as compounds that **contain nitrogen atoms bonded to carbon atoms in aliphatic chains or aromatic rings**, with the general formula **RNH_2**.

Structure

$R–NH_2$

Properties

Nitrogen has a lone pair of electrons.

Because of this,

amines can accept hydrogen ions.

That makes them **basic**.

Biological Examples

- Amino acids
- Dopamine
- Epinephrine
- Many neurotransmitters

Nursing Importance

Many drugs contain amine groups.

Their basic nature affects:

- absorption,
- distribution,
- excretion.

AMIDE

The lecture explains that in an amide, **the hydroxyl (–OH) group of a carboxylic acid is replaced by an amino ($–NH_2$) group**, forming the **$–CONH_2$** linkage.

Structure

$–CONH_2$

Why is it Important?

Amides form **peptide bonds**.

A peptide bond connects:

Amino acid

-

Amino acid

↓

Protein

Without amide bonds,

proteins would not exist.

THIOL (Mercaptan)

The lecture defines thiols as organic compounds containing a **sulfhydryl (–SH) group**, represented by the general formula **R–SH**.

Structure

R–SH

Sulfur replaces oxygen.

Biological Importance

Thiols occur in:

- cysteine,
- glutathione,
- many enzymes.

Two thiol groups can join together to form a **disulfide bond (–S–S–)**.

These bonds help stabilize protein structure.

Clinical Importance

Disulfide bonds help maintain the shape of proteins such as insulin and keratin (found in hair and nails).

ACTIVITY – The Thalidomide Tragedy

Near the end of this section, the lecture asks why the **thalidomide tragedy** happened.

The lecture places this activity just before the discussion of isomers because it introduces the importance of **molecular shape**.

The key idea is that thalidomide exists in different three-dimensional forms (isomers). Although they have the same chemical formula, they can interact differently with the body. This illustrates why the shape of a molecule is critically important in medicine. (The detailed explanation of enantiomers and stereoisomers comes in the next section of the lecture.)

⭐ Memory Tips

Functional Group	Formula	Easy Memory Trick	Example
Hydroxyl	**–OH**	**Alcohol group**	Glucose, ethanol
Ether	**R–O–R′**	Oxygen bridge	Diethyl ether
Carbonyl	**C=O**	Double-bonded oxygen	Aldehydes, ketones
Aldehyde	**R–CHO**	Carbonyl at the **end**	Glucose (open-chain)
Ketone	**R–CO–R′**	Carbonyl in the **middle**	Acetone
Carboxyl	**–COOH**	Acidic group	Amino acids, fatty acids
Carboxylic acid	**R–COOH**	Contains a carboxyl group	Acetic acid
Ester	**R–COO–R′**	From acid + alcohol	Fats, fruit aromas
Amine	**R–NH_2**	Nitrogen-containing, basic	Amino acids
Amide	**–$CONH_2$**	Peptide bond	Proteins
Thiol	**R–SH**	Sulfur-containing	Cysteine

⭐ Part 4 Summary (Must Know for Exams)

- A **functional group** is a specific arrangement of atoms attached to a carbon skeleton that determines how an organic molecule behaves and reacts.
- **Hydroxyl (–OH)** groups increase polarity and water solubility.
- **Carbonyl (C=O)** groups form the basis of **aldehydes** (end of the chain) and **ketones** (middle of the chain).
- **Carboxyl (–COOH)** groups are acidic because they can donate H^+ ions.
- **Esters** are formed from a carboxylic acid and an alcohol and are common in fats.
- **Amines (RNH_2)** are nitrogen-containing and generally basic.
- **Amides (–$CONH_2$)** form peptide bonds, which link amino acids together to make proteins.
- **Thiols (R–SH)** contain sulfur and can form disulfide bonds that stabilize protein structure.
- The transition to the **thalidomide activity** prepares you for the next topic: **isomers and molecular shape**, which explain why molecules with the same formula can have different biological effects.

PART 5 – ISOMERS AND CHIRALITY (Pages 55–65)

This is the last major topic in your lecture. It explains **isomers**, which are molecules that have the **same molecular formula but different arrangements of atoms**. Even though two isomers contain the same numbers of carbon, hydrogen, oxygen, or other atoms, their different arrangements can cause them to have **different properties and different effects in the body**. The lecture uses the **thalidomide tragedy** to show why understanding isomers is essential in medicine.

Page 55 – What Are Isomers?

The lecture defines isomers as:

Molecules that share the same molecular formula but differ in how their atoms are arranged.

Let's break that down.

Suppose two molecules both have the formula:

$\mathbf{C_4H_{10}}$

They contain:

- 4 carbon atoms
- 10 hydrogen atoms

The number of atoms is **identical**.

However, the atoms can be connected or arranged differently.

That means they are **isomers**.

Why Are Isomers Important?

Imagine two houses built with exactly:

- 1,000 bricks
- 20 windows
- 10 doors

The materials are the same.

But one house is arranged differently.

The houses will not look the same.

The same thing happens with molecules.

The atoms are the same.

The arrangement changes.

Because the arrangement changes:

- the shape changes,
- the properties change,
- the biological function may change.

This is why isomers are extremely important in biology and medicine.

Two Main Types of Isomers

The lecture divides isomers into two categories:

1. **Constitutional Isomers**
2. **Stereoisomers**

---

1. Constitutional Isomers

The lecture explains that constitutional isomers **differ in connectivity**. In other words, the atoms are bonded together in a different order.

What Does "Connectivity" Mean?

Connectivity simply means:

Which atom is attached to which atom.

Example:

Imagine four people holding hands.

If they change who they are holding hands with,

the arrangement changes.

That is exactly what happens in constitutional isomers.

Example

Butane

C–C–C–C

Isobutane

```
  C
  |
C–C–C
```

Both have:

C_4H_{10}

Same formula.

Different connections.

Therefore,

they are constitutional isomers.

Properties

Because the atoms are connected differently,

constitutional isomers often have different:

- boiling points,
- melting points,
- densities,
- chemical reactions.

---

2. Stereoisomers

The lecture defines stereoisomers as molecules with the **same chemical formula and atom connectivity but different three-dimensional spatial arrangements.**

Notice the difference.

For stereoisomers:

The atoms remain connected in the **same order**.

Only their **3D arrangement** changes.

Imagine two identical gloves.

The materials are identical.

The stitching is identical.

The only difference is that one fits the left hand and the other fits the right hand.

That is similar to stereoisomers.

Types of Stereoisomers

The lecture lists two main types:

- Geometric isomers (cis/trans)
- Optical isomers (enantiomers and diastereomers)

GEOMETRIC ISOMERS (Cis/Trans)

The lecture explains that geometric isomers are found in **alkenes and rings where rotation is restricted**.

Why Can't They Rotate?

Remember the carbon-carbon double bond?

C=C

A double bond is rigid.

It cannot rotate freely.

If it rotated,

the double bond would have to break,

which requires a lot of energy.

Because it cannot rotate,

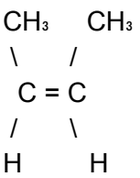
groups become fixed in position.

Cis Isomer

"Cis" means:

Same side

Example:

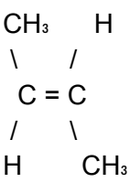


Both CH_3 groups are on the same side.

Trans Isomer

"Trans" means:

Opposite sides



The CH_3 groups are opposite each other.

Why Does This Matter?

Although both molecules contain:

- the same atoms,
- the same formula,

their shapes differ.

That difference can change:

- melting point,
- boiling point,
- biological activity.

The lecture introduces this concept as the first type of stereoisomer.

OPTICAL ISOMERS (Enantiomers)

The lecture defines enantiomers as **non-superimposable mirror images**, often containing a **chiral carbon** bonded to four different groups.

What is a Chiral Carbon?

A carbon atom becomes **chiral** when it is attached to **four different groups**.

Example:

Carbon attached to:

- H
- OH
- CH_3
- COOH

Since all four groups are different,

this carbon is chiral.

Mirror Images

Think about your hands.

Left hand

↓

Mirror

↓

Right hand

They look alike.

But they cannot perfectly overlap.

This is exactly what happens with enantiomers.

Why Are They Called Non-Superimposable?

Try placing your right hand directly on top of your left hand.

They never match perfectly.

Likewise,

one enantiomer cannot be perfectly placed on top of the other.

Why Are Enantiomers Important?

The body recognizes molecular **shape**.

Enzymes,

receptors,

and antibodies

are all shape-specific.

One enantiomer may:

- fit perfectly,
- produce the desired effect.

The other may:

- not fit,

- be less active,
- or produce unwanted effects.

This is why the correct three-dimensional arrangement matters in biology.

Diastereomers

The lecture defines diastereomers as **stereoisomers that are not mirror images of each other.**

Unlike enantiomers,

diastereomers:

- are not mirror images,
- have different physical and chemical properties.

They are another class of stereoisomers.

THE THALIDOMIDE TRAGEDY

The lecture returns to the question:

Why did the thalidomide tragedy happen?

This question connects directly to the topic of optical isomers.

The lecture's point is that **molecules with the same chemical formula can behave differently because of their three-dimensional arrangement**.

Thalidomide existed in different molecular forms (enantiomers). This illustrates that stereochemistry matters because different forms of the same compound can interact differently with biological systems. The lecture uses this historical event to emphasize the importance of understanding molecular shape in medicine.

Why Does Shape Matter in the Human Body?

Imagine an enzyme as a lock.

A molecule is the key.

Only the correctly shaped key fits.

Even if two keys are made from the same material,

a different shape means the lock may not open.

The same principle applies to:

- hormones,
- drugs,
- enzymes,
- antibodies.

Shape determines biological activity.

FINAL ACTIVITY (Page 63)

The lecture asks:

"If you had to explain this entire handout to a classmate in exactly three sentences, what would you say?"

A good answer based on the lecture would be:

1. **Life depends on organized molecules, and water and carbon provide the chemical foundation for biological structures and reactions.**
2. **Functional groups and different types of chemical bonding determine how organic molecules behave and interact inside living organisms.**
3. **The three-dimensional arrangement of atoms, especially in isomers, can greatly affect how molecules function in the body, making stereochemistry essential in biology and medicine.**

⭐ Complete Chapter 1 Summary

After studying this chapter, you should understand that:

Water

- Water is bent (104.5°) because lone pairs repel bonding pairs.
- Water is polar.
- Polarity allows hydrogen bonding.
- Hydrogen bonding explains water's high boiling point, why ice floats, and why water is an excellent solvent.

Carbon

- Carbon has four valence electrons.
- Carbon forms four covalent bonds (tetravalent).
- Carbon can form chains, branches, and rings.
- Carbon is the backbone of all major biological molecules.

Functional Groups

- Functional groups determine how organic molecules behave.
- Examples include hydroxyl, carbonyl, carboxyl, amine, amide, ester, and thiol groups.

Isomers

- Isomers have the same molecular formula but different arrangements of atoms.
- **Constitutional isomers** differ in how atoms are connected.
- **Stereoisomers** have the same connectivity but different three-dimensional arrangements.
- Stereoisomers include **geometric isomers (cis/trans)** and **optical isomers (enantiomers and diastereomers).**
- The thalidomide example highlights why molecular shape is critically important in biology and medicine.

⭐ High-Yield Exam Points

If your instructor gives a quiz on this chapter, make sure you can answer these questions:

1. Why is water polar?
2. Why does ice float?
3. Why is carbon called the backbone of life?
4. What is tetravalency?

5. What is a functional group?
6. Differentiate an aldehyde from a ketone.
7. Differentiate a carboxyl group from a carboxylic acid.
8. What is an amide bond, and why is it important in proteins?
9. What are constitutional isomers?
10. What is the difference between cis and trans isomers?
11. What is a chiral carbon?
12. Why can different isomers of the same molecule have different biological effects?

If you can confidently explain those questions, you'll have covered the core learning objectives from this chapter.