# Chapter 5: Chemical Bonding

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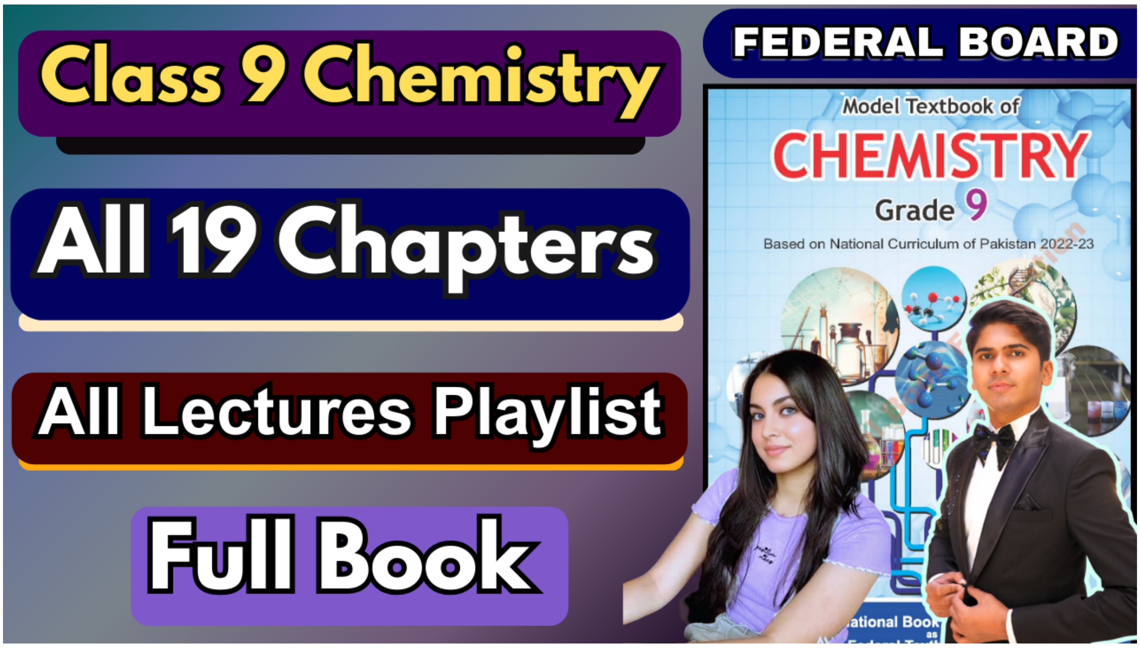


Most matter in the world is composed of compounds and their mixtures. Examples include human/animal/plant bodies, rocks, soil, petroleum, and coal. Compounds are formed when different kinds of atoms are bonded together. A few elements (e.g., Noble Gases like Helium, Neon, Argon) exist as unbonded atoms. The properties of a substance (hardness, flexibility, stickiness) are directly determined by the nature of the bonding and the structure of its molecules.

## 5.1. Why Do Atoms React?

### The Octet and Duplet Rules (G.N. Lewis, 1916)

These rules explain the reactivity and stability of atoms by focusing on their valence electrons.

#### The Octet Rule:

An atom is most stable when its valence shell contains **eight electrons**. It applies to main group elements and involves only $s$ and $p$ electrons. Examples: Oxygen, Nitrogen, and Carbon follow this rule. Atoms gain, lose, or share electrons to achieve a full octet.

- **Sodium (Na):** Unstable with an incomplete octet. Loses 1 electron to form $Na^+$, which has the same electron configuration as Neon (Ne).
- **Chlorine (Cl):** Unstable with an incomplete octet. Gains 1 electron to form $Cl^-$, which has the same electron configuration as Argon (Ar).

#### The Duplet Rule:

The tendency of atoms to acquire a **two-electron** configuration (like Helium) in their outermost shell. Applies to elements whose valence electrons are only in the $s$ orbital (e.g., Hydrogen, Lithium, Beryllium).

## 5.2. Ionic Bonding

Formed by the complete **transfer of electrons** from a metal atom to a non-metal atom. This transfer creates ions (charged atoms).

- **Metal** (low Ionization Energy) $\rightarrow$ **Loses** electrons $\rightarrow$ forms **Cation** (positive ion).
- **Non-metal** (high Electron Affinity) $\rightarrow$ **Gains** electrons $\rightarrow$ forms **Anion** (negative ion).

The bond is the strong electrostatic force of attraction between the opposite charges of the cation and anion.

#### Example: Formation of Sodium Chloride (NaCl)

- Na (2, 8, 1) loses 1 electron to become stable $Na^+$ (2, 8).
- Cl (2, 8, 7) gains 1 electron to become stable $Cl^-$ (2, 8, 8).
- The electrostatic attraction between $Na^+$ and $Cl^-$ forms NaCl.

#### Properties of Ionic Compounds

- **High Melting and Boiling Points:** Due to the strong forces holding the ions together in a crystal lattice. A lot of energy is needed to break this lattice.
- **Conductivity:**
  - **Solid State:** Do not conduct electricity because ions are fixed in their positions.
  - **Molten (Liquid) or Aqueous (Dissolved in water) State:** Conduct electricity because the ions become free to move and carry the charge.
- **Hard and Brittle:** They are hard because of strong forces, but brittle because a small shift in the lattice brings like-charged ions next to each other, causing strong repulsion and fracturing the crystal.

## 5.3. Covalent Bonding

Formed by the **mutual sharing of electrons** between two non-metal atoms (or sometimes a non-metal and a metalloid). Both atoms contribute to the shared electrons to complete their Octet (or Duplet).

#### Types of Covalent Bonds

- **Single Bond:** Sharing of **one pair** of electrons (e.g., $H_2, Cl_2$)
- **Double Bond:** Sharing of **two pairs** of electrons (e.g., $O_2$)
- **Triple Bond:** Sharing of **three pairs** of electrons (e.g., $N_2$)

### Polar and Non-Polar Covalent Bonds

- **Non-Polar Covalent Bond:** Electrons are **shared equally** between two atoms (Electronegativity difference is zero or very small). The charge is evenly distributed. E.g., $H_2, Cl_2$
- **Polar Covalent Bond:** Electrons are **shared unequally** (Electronegativity difference is $0.4$ to $1.7$). The atom with higher electronegativity pulls the shared electrons closer, creating a **dipole** (partial charges). E.g., $HCl, H_2O$

### Coordinate Covalent Bond (Dative Bond)

A special type of covalent bond where **both shared electrons** in the bond are donated by **only one atom**. The donor atom must have a **lone pair** (a pair of non-bonded valence electrons). The acceptor atom must have an empty orbital. Example: Formation of Ammonium ion ($NH_4^+$) from $NH_3$ and $H^+$

### Properties of Covalent Compounds

- **Low Melting and Boiling Points:** Most covalent compounds exist as discrete molecules held by weak intermolecular forces (forces between molecules), which are easy to break.
- **Conductivity:** Usually **poor conductors** of electricity (in all states) because they do not form free ions.
- **Solubility:**
  - **Non-polar** compounds (like oil) dissolve in **non-polar** solvents (like petrol).
  - **Polar** compounds (like sugar) dissolve in **polar** solvents (like water).

## 5.4. Hydrogen Bonding (Intermolecular Force)

A **strong type of dipole-dipole attraction** that occurs when a hydrogen atom (H) is bonded directly to a small, highly electronegative atom like **Fluorine (F), Oxygen (O), or Nitrogen (N)**. This creates a very strong partial positive charge on the hydrogen atom, which is then strongly attracted to the lone pair of electrons on a nearby F, O, or N atom of another molecule.

### Effects of Hydrogen Bonding

- **High Boiling Point:** Substances like $H_2O$, HF, and $NH_3$ have unusually high boiling points compared to similar compounds without H-bonding, because extra energy is required to break these strong intermolecular forces.
- **Low Density of Ice:** In ice, H-bonds create an open, cage-like structure that makes solid water less dense than liquid water, which is why ice floats.

## 5.5. Allotropes and Allotropy

The existence of an element in **two or more different physical forms** in the same state (solid, liquid, or gas) is called allotropy. These forms (allotropes) have different physical properties but the same chemical properties.

### 5.5.1. Graphite

A form of carbon in which atoms are arranged in **flat, hexagonal layers**.

- Each carbon atom is bonded to only **three** others in a layer.
- The layers are held together by **weak van der Waals forces** and can easily slide over each other.

**Why it's useful (because of its structure):**

- **Soft and Slippery:** Used as a **lubricant** and in **pencil lead** because the layers can slide easily.
- **Conductor of Electricity:** Contains **free/delocalized electrons** that can move within the layers to conduct electricity (unlike diamond).
- **Refractory Material:** Used in crucibles and furnace linings due to its high melting point and resistance to heat.
- **Moderator in Nuclear Reactors:** Slows down fast-moving neutrons.

### 5.5.2. Diamond

Diamond is the **hardest natural material** known. It is a brilliant, transparent form of carbon where the atoms are locked into an incredibly strong **3D network**. Each carbon atom is strongly **covalently bonded to four other carbon atoms** in a rigid, tetrahedral shape. This creates a giant, three-dimensional lattice that extends in every direction. Breaking a diamond means breaking these incredibly strong bonds. **Why it's useful (because of its structure):**

- **Extreme Hardness:** Its rigid 3D structure makes it the hardest material. It's used on the tips of drill bits, cutting tools, and grinding wheels to cut through other hard materials.
- **Brilliance & Luster:** After being cut and polished, diamonds brilliantly reflect light, making them prized for jewellery.
- **Other Uses:** Its properties also make it useful in specialized medical equipment, like surgical tools, and in high-quality sound equipment for DJs because it can vibrate very quickly without distorting.

## 5.6. Metallic Bonding (e.g., Fe, Cu, Al)

A "**sea of delocalized electrons**" surrounding a lattice of positive metal ions. The attraction between the electrons and ions is the **metallic bond**.

#### Properties (Explained by the Electron Sea Model)

- **Malleable/Ductile:** Layers of ions can slide without breaking because the electron sea holds everything together.
- **Good Conductors:** The delocalized electrons are free to move and carry charge/heat.
- **High MP/BP:** Strong metallic bonds.


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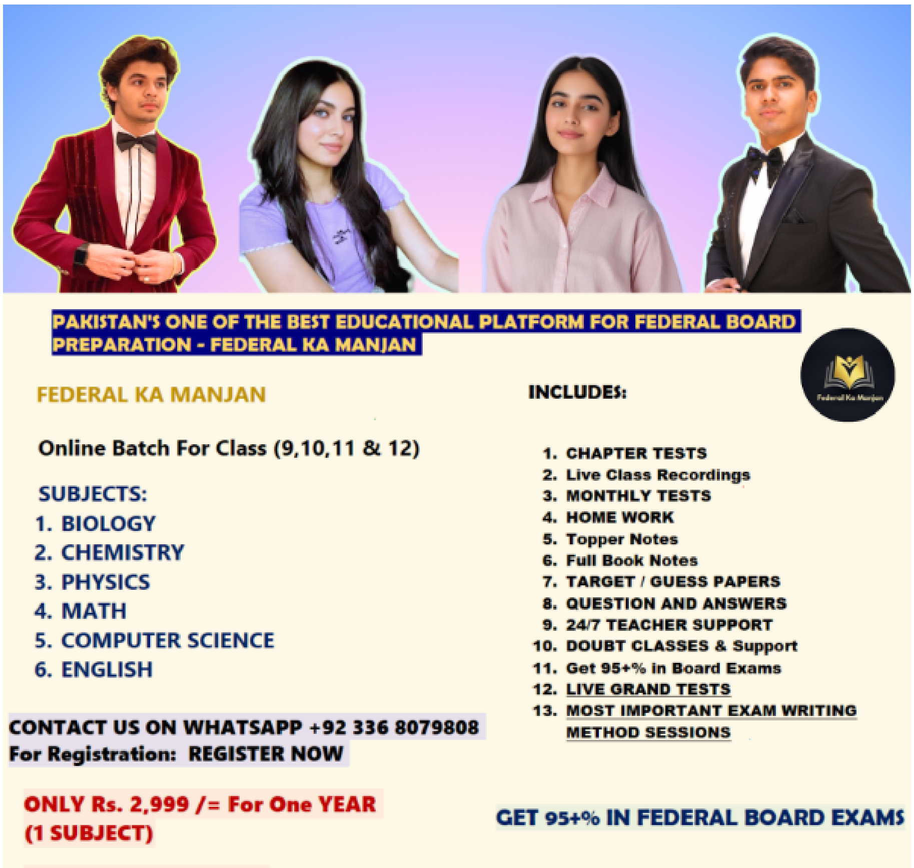
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