# Unit 2: Chemical Bonding

- **Chemical bonds are the attractive forces that hold atoms, ions, and molecules together.** These bonds arise from the interactions between the electrons of participating atoms, particularly their valence electrons, which are the electrons in the outermost occupied energy level. As you learned in Unit 1, valence electrons are responsible for an atom's chemical reactivity.
- **One of the fundamental principles guiding our understanding of chemical bonding is the octet rule.** This rule states that atoms tend to gain, lose, or share electrons to achieve a stable electron configuration with eight electrons in their outermost shell, resembling the stable electron configuration of noble gases. The octet rule helps to explain why certain elements form specific types of bonds and adopt predictable charges in their ionic forms.
- **There are three primary types of chemical bonds, each with its own characteristics and properties:**
    1. **Ionic bonds** are formed between a metal and a non-metal. In ionic bonding, electrons are transferred from the metal atom to the non-metal atom, creating oppositely charged ions that attract each other strongly. This electrostatic attraction between cations (positive ions) and anions (negative ions) forms the basis of the ionic bond.
        - For example, in the formation of sodium chloride (NaCl), sodium (Na), a metal, readily loses its one valence electron to achieve a stable octet, forming a $Na^+$ cation. Chlorine (Cl), a non-metal with seven valence electrons, readily accepts this electron to complete its octet, becoming a $Cl^-$ anion. The electrostatic attraction between $Na^+$ and $Cl^-$ forms the ionic bond in NaCl.
        - **Ionic compounds** tend to form crystalline structures, have high melting and boiling points, and are good conductors of electricity when molten or dissolved in water due to the presence of mobile ions.
    2. **Covalent bonds** are formed between two non-metals. In covalent bonding, atoms share electrons to achieve a stable octet. The shared electrons are attracted to the nuclei of both bonding atoms, creating a force that holds the atoms together.
        - For example, in a molecule of hydrogen ($H_2$), each hydrogen atom shares its single electron with the other, effectively giving both atoms a stable duet (two electrons) resembling the electron configuration of the noble gas helium.
        - **Covalent bonds can be single, double, or triple, depending on the number of electron pairs shared between the atoms.**
        - **Covalent molecules can be polar or non polar depending on the electronegativity difference between the bonding atoms and the molecular geometry.** In a **polar covalent bond**, the electrons are shared unequally, resulting in a partial positive charge ($\delta+$) on one atom and a partial negative charge ($\delta-$) on the other.

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        - **Covalent compounds** can exist as solids, liquids, or gases at room temperature, have relatively low melting and boiling points compared to ionic compounds, and are generally poor conductors of electricity.
    3. **Metallic bonds** are formed between metal atoms. In metallic bonding, valence electrons are delocalised, meaning they are free to move throughout the entire metal lattice. This creates a "sea of electrons" that surrounds and binds together the positively charged metal ions. This unique bonding model accounts for many of the characteristic properties of metals, such as:
        - **Good electrical conductivity:** The delocalised electrons are free to move and carry an electrical current.
        - **Malleability (ability to be hammered into sheets) and ductility (ability to be drawn into wires):** The metal ions can slide past each other without disrupting the metallic bonding due to the delocalised electron sea.
- **In addition to intramolecular forces (ionic, covalent, and metallic bonds), intermolecular forces also play a significant role in the properties of matter.** These forces are weaker than the primary chemical bonds but influence the physical properties of substances.

    - **Hydrogen bonding**, a special type of dipole-dipole interaction, is the strongest of these intermolecular forces. It occurs when a hydrogen atom is bonded to a highly electronegative atom (such as oxygen, nitrogen, or fluorine) and is attracted to a lone pair of electrons on a neighbouring molecule. Hydrogen bonds are crucial in biological systems, influencing the structure and properties of water, proteins, and DNA.
    - **Dipole-dipole forces** exist between polar molecules, arising from the attraction between the positive end (δ+) of one molecule and the negative end (δ-) of another.
    - **London dispersion forces**, the weakest of the intermolecular forces, are present in all substances, even non polar molecules. These forces arise due to temporary fluctuations in electron distribution within molecules, creating temporary dipoles that induce opposite dipoles in neighbouring molecules.
- **The Valence Bond Theory (VBT) and Molecular Orbital Theory (MOT) provide more sophisticated models for understanding covalent bonding.**
    - **VBT** describes covalent bond formation as the overlapping of atomic orbitals between two atoms, with the shared electron pair occupying the region of overlap. It introduces the concept of **hybridization**, where atomic orbitals mix to form hybrid orbitals with specific geometries that influence molecular shapes. For example, sp3 hybridization leads to a tetrahedral geometry, explaining the shape of molecules like methane (CH4).
    - **MOT** takes a different approach, considering the combination of atomic orbitals to form molecular orbitals that extend over the entire molecule. It introduces bonding and antibonding orbitals and helps explain molecular properties like bond order and magnetic behaviour. MOT is particularly useful in understanding molecules that VBT cannot adequately explain.

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