Notes for Exam Preparation: : Chemical Bonding

Introduction

Chemical bonding is the force that holds atoms together in compounds. Understanding the nature of these bonds is crucial for explaining the physical and chemical properties of substances. There are three main types of chemical bonding: ionic, covalent, and metallic. In addition, intermolecular forces like van der Waals' forces, permanent dipole-dipole interactions, and hydrogen bonding play significant roles in determining the behavior of molecules.

Types of Chemical Bonding

Van der Waals' Forces

Van der Waals' forces, also known as dispersion forces or temporary dipole–induced dipole forces, are weak interactions that occur between all atoms and molecules due to the temporary distribution of electrons.

- **Temporary Dipole–Induced Dipole Forces**: These arise when the random movement of electrons creates a temporary dipole in one molecule, which then induces a dipole in a neighboring molecule.
- **Strength**: These forces are generally weak but increase with the size and mass of the molecules.

Permanent Dipole–Dipole Forces

Permanent dipole–dipole forces occur between molecules that have permanent dipoles. These interactions are stronger than van der Waals' forces and occur in polar molecules where there is a permanent separation of charge.

- **Example**: Hydrogen chloride (HCl), where the positive end (H) of one molecule is attracted to the negative end (Cl) of another.

Hydrogen Bonds

Hydrogen bonds are a special type of dipole-dipole interaction that occur when hydrogen is bonded to highly electronegative atoms like nitrogen, oxygen, or fluorine.

- **Strength**: Hydrogen bonds are stronger than other dipole-dipole interactions but weaker than covalent and ionic bonds.
- **Examples**: Water (H_2O), ammonia (NH_3), and hydrogen fluoride (HF).

Ionic Bonding

Ionic bonding involves the transfer of electrons from one atom to another, resulting in the formation of positively and negatively charged ions.

How are Ions Formed?

- **Cations**: Formed when an atom loses one or more electrons.
- **Anions**: Formed when an atom gains one or more electrons.

Dot-and-Cross Diagrams

Dot-and-cross diagrams are used to represent the transfer of electrons in ionic bonding.

Examples

- **Magnesium Oxide (MgO)**:
  - Magnesium (Mg) loses two electrons to form Mg2+Mg^{2+}Mg2+.
  - Oxygen (O) gains two electrons to form O2−O^{2-}O2−.
- Mg: · · · · · · O: · · · · · · \begin{array}{ccc} \text{Mg:} & \cdot \cdot & \cdot \cdot \cdot \cdot \\ \text{O:} & \cdot \cdot \cdot \cdot & \cdot \cdot \\ \end{array}Mg:O: · · · · · · · · · · · ·
- **Calcium Chloride (CaCl2)**:
  - Calcium (Ca) loses two electrons to form Ca2+Ca^{2+}Ca2+.
  - Each chlorine (Cl) gains one electron to form two Cl−Cl^{-}Cl−.
- Ca: · · · · Cl: · · · · · · \begin{array}{ccc} \text{Ca:} & \cdot \cdot & \cdot \cdot \\ \text{Cl:} & \cdot \cdot \cdot \cdot & \cdot \cdot \\ \end{array}Ca:Cl: · · · · · · · · · ·

Covalent Bonding

Covalent bonding involves the sharing of electron pairs between atoms.

Single Covalent Bonds

A single covalent bond consists of one shared pair of electrons.

- **Lone Pairs**: Electrons that are not involved in bonding.

Multiple Covalent Bonds

- **Double Bonds**: Involve two shared pairs of electrons (e.g., O2).
- **Triple Bonds**: Involve three shared pairs of electrons (e.g., N2).

Co-ordinate Bonding (Dative Covalent Bonding)

A co-ordinate bond occurs when both electrons in the shared pair come from the same atom.

- **Example**: Ammonium ion (NH_4^+), where nitrogen donates a lone pair to bond with a hydrogen ion (H^+).

Bond Length and Bond Energy

- **Bond Length**: The distance between the nuclei of two bonded atoms.
- **Bond Energy**: The energy required to break one mole of a bond in a molecule in the gas phase.

Shapes of Molecules

The shape of a molecule is determined by the number of electron pairs (bonding and lone pairs) around the central atom.

Electron-Pair Repulsion Theory

Electron pairs repel each other and arrange themselves as far apart as possible around the central atom.

Working Out the Shapes of Molecules

- **Linear**: 2 bonding pairs, 0 lone pairs (e.g., CO_2).
- **Trigonal Planar**: 3 bonding pairs, 0 lone pairs (e.g., BF_3).
- **Tetrahedral**: 4 bonding pairs, 0 lone pairs (e.g., CH_4).
- **Trigonal Bipyramidal**: 5 bonding pairs, 0 lone pairs (e.g., PCl_5).
- **Octahedral**: 6 bonding pairs, 0 lone pairs (e.g., SF_6).
- **Bent/V-Shaped**: 2 bonding pairs, 1 or 2 lone pairs (e.g., H_2O).
- **Trigonal Pyramidal**: 3 bonding pairs, 1 lone pair (e.g., NH_3).

Summary

Chemical bonding is a fundamental concept in chemistry that explains the interactions between atoms and the formation of compounds. Ionic, covalent, and metallic bonds are the primary types of chemical bonds, each with unique characteristics and behaviors. Intermolecular forces, such as van der Waals' forces, dipole-dipole interactions, and hydrogen bonds, also play critical roles in determining the properties of substances. Understanding these concepts is essential for predicting and explaining the properties and behaviors of different materials.

Practice Questions and Answers

Question 1:

Draw a dot-and-cross diagram for magnesium oxide (MgO).

Answer:

- Magnesium loses two electrons to form Mg2+Mg^{2+}Mg2+.
- Oxygen gains two electrons to form O2−O^{2-}O2−.
  Mg:· ·→[Mg2+]O:· · · ·←· ·[O2−]\text{Mg:} \cdot \cdot \rightarrow [\text{Mg}^{2+}] \\ \text{O:} \cdot \cdot \cdot \cdot \leftarrow \cdot \cdot [\text{O}^{2-}]Mg:· ·→[Mg2+]O:· · · ·←· ·[O2−]

Question 2:

Explain the difference between ionic and covalent bonding.

Answer:

- **Ionic Bonding**: Involves the transfer of electrons from one atom to another, resulting in the formation of positive and negative ions.
- **Covalent Bonding**: Involves the sharing of electron pairs between atoms to achieve a stable electron configuration.

Question 3:

Predict the shape of a methane (CH4) molecule and explain your reasoning.

Answer: Methane (CH4) has a tetrahedral shape. This is because the central carbon atom has four bonding pairs of electrons that repel each other to the maximum extent, arranging themselves as far apart as possible in a tetrahedral geometry.

Types of Chemical Bonding

Van der Waals' Forces

Van der Waals' forces, also known as dispersion forces or temporary dipole–induced dipole forces, are weak interactions that occur between all atoms and molecules due to temporary distributions of electrons.

- **Temporary Dipole–Induced Dipole Forces**: These arise when the random movement of electrons creates a temporary dipole in one molecule, which then induces a dipole in a neighboring molecule.
- **Strength**: Generally weak but increase with the size and mass of the molecules.

Permanent Dipole–Dipole Forces

Permanent dipole–dipole forces occur between molecules that have permanent dipoles. These interactions are stronger than van der Waals' forces and occur in polar molecules where there is a permanent separation of charge.

- **Example**: Hydrogen chloride (HCl), where the positive end (H) of one molecule is attracted to the negative end (Cl) of another.

Hydrogen Bonds

Hydrogen bonds are a special type of dipole-dipole interaction that occurs when hydrogen is bonded to highly electronegative atoms like nitrogen, oxygen, or fluorine.

- **Strength**: Hydrogen bonds are stronger than other dipole-dipole interactions but weaker than covalent and ionic bonds.
- **Examples**: Water (H_2O), ammonia (NH_3), and hydrogen fluoride (HF).

Ionic Bonding

Ionic bonding involves the transfer of electrons from one atom to another, resulting in the formation of positively and negatively charged ions.

How are Ions Formed?

- **Cations**: Formed when an atom loses one or more electrons.
- **Anions**: Formed when an atom gains one or more electrons.

Dot-and-Cross Diagrams

Dot-and-cross diagrams represent the transfer of electrons in ionic bonding.

Examples:

- **Magnesium Oxide (MgO)**:
  - Magnesium (Mg) loses two electrons to form Mg2+Mg^{2+}Mg2+.
  - Oxygen (O) gains two electrons to form O2−O^{2-}O2−.
- **Calcium Chloride ($CaCl_2$)**:
  - Calcium (Ca) loses two electrons to form Ca2+Ca^{2+}Ca2+.
  - Each chlorine (Cl) gains one electron to form Cl−Cl^{-}Cl−.

Covalent Bonding

Covalent bonding involves the sharing of electron pairs between atoms.

Single Covalent Bonds

A single covalent bond consists of one shared pair of electrons.

- **Lone Pairs**: Electrons that are not involved in bonding.

Multiple Covalent Bonds

- **Double Bonds**: Involve two shared pairs of electrons (e.g., O_2).
- **Triple Bonds**: Involve three shared pairs of electrons (e.g., N_2).

Co-ordinate Bonding (Dative Covalent Bonding)

A co-ordinate bond occurs when both electrons in the shared pair come from the same atom.

- **Example**: Ammonium ion (NH_4^+), where nitrogen donates a lone pair to bond with a hydrogen ion (H^+).

Bond Length and Bond Energy

- **Bond Length**: The distance between the nuclei of two bonded atoms.
- **Bond Energy**: The energy required to break one mole of a bond in a molecule in the gas phase.

Shapes of Molecules

The shape of a molecule is determined by the number of electron pairs (bonding and lone pairs) around the central atom.

Electron-Pair Repulsion Theory

Electron pairs repel each other and arrange themselves as far apart as possible around the central atom.

Working Out the Shapes of Molecules

- **Linear**: 2 bonding pairs, 0 lone pairs (e.g., CO_2).
- **Trigonal Planar**: 3 bonding pairs, 0 lone pairs (e.g., BF_3).
- **Tetrahedral**: 4 bonding pairs, 0 lone pairs (e.g., CH_4).
- **Trigonal Bipyramidal**: 5 bonding pairs, 0 lone pairs (e.g., PCl_5).
- **Octahedral**: 6 bonding pairs, 0 lone pairs (e.g., SF_6).
- **Bent/V-Shaped**: 2 bonding pairs, 1 or 2 lone pairs (e.g., H_2O).
- **Trigonal Pyramidal**: 3 bonding pairs, 1 lone pair (e.g., NH_3).

Shapes of Some Organic Molecules

- **Ethane (C_2H_6)**: Tetrahedral around each carbon atom.
- **Ethene (C_2H_4)**: Trigonal planar around each carbon atom due to double bond.
- **Ethyne (C_2H_2)**: Linear around each carbon atom due to triple bond.

Sigma and Pi Bonds

Sigma (σ) Bonds

- **Formation**: Formed by the end-to-end overlap of orbitals.
- **Characteristics**: Stronger and allow free rotation around the bond axis.

Pi (π) Bonds

- **Formation**: Formed by the side-to-side overlap of p-orbitals.
- **Characteristics**: Weaker than sigma bonds and restrict rotation around the bond axis.

Metallic Bonding

What is a Metallic Bond?

A metallic bond is the attraction between positively charged metal ions and the sea of delocalized electrons.

Delocalized Electrons

- **Characteristics**: Electrons are free to move throughout the metal lattice, providing conductivity.

Strength of Metallic Bonding

The strength of metallic bonding increases with:

- Increasing positive charge on the ions in the metal lattice.
- Decreasing size of metal ions in the lattice.
- Increasing number of mobile electrons per atom.

Properties of Metals

- **High Melting and Boiling Points**: Due to strong metallic bonds.
- **Electrical Conductivity**: Due to delocalized electrons.
- **Thermal Conductivity**: Also due to delocalized electrons.

Intermolecular Forces

Types of Intermolecular Forces

- **Van der Waals' Forces**: Weak forces that increase with the number of electrons and contact points between molecules.
- **Permanent Dipole–Dipole Forces**: Occur between polar molecules.
- **Hydrogen Bonds**: Strong dipole-dipole interactions involving hydrogen and electronegative atoms.

Effect on Melting and Boiling Points

- **Van der Waals' Forces**: Increase with molecular size.
- **Permanent Dipole–Dipole Forces**: Increase melting and boiling points of polar substances.
- **Hydrogen Bonds**: Significantly increase boiling points (e.g., water).

Polarity in Molecules

Electronegativity

- **Pattern**: Increases across a period and decreases down a group in the periodic table.
- **Polarity**: Difference in electronegativity between atoms in a bond creates a dipole.

Polarity and Chemical Reactivity

- **Polar Molecules**: Reactivity influenced by dipole moments, affecting interactions with other polar or ionic substances.

Bonding and Physical Properties

Physical State at Room Temperature and Pressure

- **Ionic Compounds**: Usually solid due to strong electrostatic forces.
- **Metals**: Solid with high melting points due to metallic bonds.
- **Covalent Compounds**: Can be gases, liquids, or solids depending on intermolecular forces.

Solubility

- **Ionic Compounds**: Soluble in water due to ion-dipole interactions.
- **Metals**: Generally insoluble in water, though reactive metals may react with water.
- **Covalent Compounds**: Solubility varies with polarity.

Electrical Conductivity

- **Ionic Compounds**: Conduct electricity when molten or in solution.
- **Metals**: Conduct electricity due to free-moving electrons.
- **Covalent Compounds**: Generally do not conduct electricity (except graphite).

Summary

Understanding the various types of chemical bonding and the resulting molecular shapes is crucial for predicting the properties and behaviors of substances. Intermolecular forces, metallic bonding, and the concept of polarity play significant roles in determining the physical states, solubility, and conductivity of compounds.

Practice Questions and Answers

Question 1:

Draw a dot-and-cross diagram for calcium chloride (CaCl2).

Answer:

- Calcium loses two electrons to form Ca2+Ca^{2+}Ca2+.
- Each chlorine gains one electron to form Cl−Cl^{-}Cl−.

Question 2:

Predict the shape of a phosphorus pentachloride (PCl5) molecule and explain your reasoning.

Answer: Phosphorus pentachloride (PCl5) has a trigonal bipyramidal shape. This is because the central phosphorus atom has five bonding pairs of electrons that repel each other to the maximum extent, arranging themselves in a trigonal bipyramidal geometry.

Question 3:

Explain the difference between sigma and pi bonds.

Answer:

- **Sigma (σ) Bonds**: Formed by the end-to-end overlap of orbitals, are stronger and allow free rotation around the bond axis.
- **Pi (π) Bonds**: Formed by the side-to-side overlap of p-orbitals, are weaker and restrict rotation around the bond axis.

Question 4:

How does hydrogen bonding affect the boiling point of water?

Answer: Hydrogen bonding significantly increases the boiling point of water because the strong hydrogen bonds require more energy to break compared to other dipole-dipole interactions..