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CHM101

GENERAL CHEMISTRY I

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Course code: CHM 101

Course Title: General Chemistry I

Course Unit: 2 Units

Series 1: Foundations of Atoms, Molecules, and Chemical Reactions

Part 1 Atoms and Molecules: The Building Blocks of Matter

- Definition and concept of atoms
- Molecules and compounds
- Elements and symbols

Part 2 Chemical Reactions and Equations

- Nature of chemical reactions
- Types of chemical reactions (combination, decomposition, displacement, exchange)
- Writing and balancing chemical equations

Part 3 Stoichiometry and Quantitative Relationships

- The mole concept
- Mass, volume, and particle relationships in reactions
- Applications of stoichiometry in problem solving

Part 4 Introduction to Redox Reactions

- Oxidation and reduction definitions
- Identifying oxidising and reducing agents
- Balancing simple redox equations

Introduction

Chemistry begins with the smallest units of matter that make up everything around us. From the air we breathe to the food we eat, all substances are composed of atoms and molecules that interact through chemical reactions. Understanding these fundamental concepts provides the foundation upon which the rest of chemistry is built. This Series will therefore introduce you to the essential principles that explain how matter is structured and how it changes, preparing you for deeper exploration in subsequent Series.

The aim of this Series is to equip you with the knowledge and skills needed to define atoms and molecules, describe their interactions, and apply the principles of chemical equations and stoichiometry to real problems. By the end of this Series, you will be able to explain the basic building blocks of matter and demonstrate how chemical reactions are represented and analysed.

We shall commence this Series by examining atoms and molecules as the basic building blocks of matter. Next, we will move on to chemical reactions and equations, where you will learn how to classify, write, and balance them. Following that, we will explore stoichiometry and quantitative relationships, focusing on the mole concept and its applications in problem solving. Finally, we will wrap up by introducing redox reactions, where you will identify oxidation and reduction processes and practise balancing simple redox equations.

Series Learning Outcomes

Upon completing this Series, you should be able to:

- **SLO 1.1** define atoms and molecules as the fundamental building blocks of matter,
- **SLO 1.2** explain the distinction between elements, compounds, and their symbolic representations,
- **SLO 1.3** describe the nature and types of chemical reactions,
- **SLO 1.4** demonstrate how to write and balance chemical equations accurately,
- **SLO 1.5** apply the mole concept to calculate mass, volume, and particle relationships in chemical reactions,
- **SLO 1.6** analyse stoichiometric problems to determine quantitative relationships in reactions, and
- **SLO 1.7** identify oxidation and reduction processes and balance simple redox equations.

1.1 Atoms and Molecules: The Building Blocks of Matter

1.1.1 Definition and concept of atoms

An **atom** is the smallest unit of an element that retains the chemical properties of that element. It consists of a nucleus, which contains positively charged protons and neutral neutrons, surrounded by negatively charged electrons that occupy specific energy levels. Atoms are the fundamental units of matter and combine in various ways to form molecules and compounds.

Atoms are incredibly small, yet they form the basis of all substances. For example, the water you drink is made up of atoms of hydrogen and oxygen bonded together.

Fill in the blank: The smallest unit of an element that retains its chemical properties is called an _____.

1.1.2 Molecules and compounds

A **molecule** is a group of two or more atoms held together by chemical bonds. Molecules may consist of atoms of the same element, such as oxygen gas (O_2), or atoms of different elements, such as water (H_2O).

A **compound** is a substance formed when atoms of different elements combine in fixed proportions. Compounds have properties distinct from the individual elements that make them up. For instance, sodium is a reactive metal and chlorine is a poisonous gas, yet together they form sodium chloride (NaCl), which is common table salt.

Fill in the blank: A substance formed when atoms of different elements combine in fixed proportions is called a _____.

1.1.3 Elements and symbols

An **element** is a pure substance made up of only one type of atom. Each element is represented by a unique chemical symbol, usually derived from its English or Latin name. For example, hydrogen is represented by H, oxygen by O, and sodium by Na (from the Latin word *natrium*).

The periodic table provides a systematic arrangement of all known elements, each with its symbol, atomic number, and other properties. Learning these symbols is essential for writing chemical equations and understanding chemical reactions.

Fill in the blank: The chemical symbol for sodium is derived from its Latin name, _____.

Element	Symbol
Hydrogen	H
Carbon	C
Nitrogen	N
Oxygen	O
Sodium	Na
Magnesium	Mg
Aluminium	Al
Silicon	Si
Phosphorus	P
Sulfur	S
Chlorine	Cl
Potassium	K
Calcium	Ca
Iron	Fe
Copper	Cu
Zinc	Zn
Silver	Ag
Gold	Au
Mercury	Hg
Lead	Pb

Table 1.1: Selected elements and their symbols

Pilot questions 1.1

1. Which subatomic particles are found in the nucleus of an atom?
2. What is the smallest unit of matter that retains the properties of an element?
3. Give one example of a molecule made up of atoms of the same element.
4. What is the difference between a molecule and a compound?
5. Write the chemical symbol for chlorine.

1.2 Chemical Reactions and Equations

1.2.1 Nature of chemical reactions

A **chemical reaction** is a process in which one or more substances, called reactants, are transformed into new substances, called products. During a chemical reaction, bonds between atoms are broken and new bonds are formed, resulting in changes in composition and properties. For example, when hydrogen reacts with oxygen, water is formed.

Chemical reactions are central to chemistry because they explain how substances interact and change. They can involve energy changes, such as the release of heat in combustion or the absorption of energy in photosynthesis.

Fill in the blank: Substances that undergo change in a chemical reaction are called _____.

1.2.2 Types of chemical reactions

Chemical reactions can be classified into several types:

- **Combination reaction**: Two or more substances combine to form a single product. Example: $2H_2 + O_2 \rightarrow 2H_2O$.
- **Decomposition reaction**: A single compound breaks down into two or more simpler substances. Example: $2HgO \rightarrow 2Hg + O_2$.
- **Displacement reaction**: One element replaces another in a compound. Example: $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$.
- **Exchange (double displacement) reaction**: Two compounds exchange ions to form new compounds. Example: $NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$.

Fill in the blank: A reaction in which a single compound breaks down into simpler substances is called a _____ reaction.

[Insert box placeholder here] Caption: *Box 1.1: Summary of types of chemical reactions*

- Box listing combination, decomposition, displacement, and exchange reactions with examples
- Suggested AI prompt: “Create a summary box listing four types of chemical reactions with examples.”
- Search string: “OER summary types of chemical reactions examples”

1.2.3 Writing and balancing chemical equations

A **chemical equation** is a symbolic representation of a chemical reaction, showing the reactants, products, and their relative amounts. Balancing chemical equations ensures that the law of conservation of mass is obeyed, meaning the number of atoms of each element is the same on both sides of the equation.

To balance an equation, coefficients are placed before chemical formulas to adjust the number of atoms. For example, the combustion of methane is written as: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.

Balancing equations requires practice and careful attention to detail. Always begin with the most complex molecule and adjust coefficients systematically.

Fill in the blank: The principle that requires the number of atoms of each element to be equal on both sides of a chemical equation is called the law of _____.

Pilot questions 1.2

1. What is the difference between reactants and products in a chemical reaction?
2. Write one example of a combination reaction.
3. Which type of reaction involves the exchange of ions between two compounds?
4. Why must chemical equations be balanced?
5. Balance the equation: $H_2 + O_2 \rightarrow H_2O$.

1.3 Stoichiometry and Quantitative Relationships

1.3.1 The mole concept

The **mole** is a fundamental unit in chemistry used to measure the amount of substance. One mole contains 6.022×1023particles, which may be atoms, molecules, or ions. This number is known as **Avogadro’s constant**. The mole allows chemists to relate the microscopic world of atoms and molecules to measurable quantities in the laboratory.

For example, one mole of water molecules contains 6.022×1023molecules of H_2O, and its mass is 18 grams.

Fill in the blank: The number of particles in one mole of a substance is called _____ constant.

1.3.2 Mass, volume, and particle relationships in reactions

Stoichiometry uses the mole concept to calculate relationships between reactants and products in a chemical reaction. These relationships can be expressed in terms of **mass**, **volume**, or **number of particles**.

- **Mass relationships**: Based on molar masses of substances. For example, in the reaction $2H_2 + O_2 \rightarrow 2H_2O$, 4 g of hydrogen reacts with 32 g of oxygen to produce 36 g of water.
- **Volume relationships**: For gases, volumes can be related using the molar volume at standard temperature and pressure (STP), which is 22.4 litres per mole.
- **Particle relationships**: The number of atoms or molecules can be calculated using Avogadro's constant.

  Fill in the blank: At STP, one mole of any gas occupies _____ litres.

1.3.3 Applications of stoichiometry in problem solving

Stoichiometry is applied to predict the amounts of reactants required or products formed in a chemical reaction. It is essential for laboratory work, industrial processes, and environmental studies.

For example, in combustion reactions, stoichiometry helps determine the exact amount of oxygen needed to burn a given fuel completely. In titration experiments, it allows chemists to calculate concentrations of unknown solutions.

Fill in the blank: Stoichiometry ensures that chemical reactions are carried out with the correct _____ of reactants and products.

Pilot questions 1.3

1. What is the value of Avogadro's constant?
2. How many molecules are present in one mole of water?
3. At STP, what volume does one mole of gas occupy?
4. In the reaction $2H_2 + O_2 \rightarrow 2H_2O$, how many grams of oxygen are required to react with 4 g of hydrogen?
5. Why is stoichiometry important in chemical problem solving?

1.4 Introduction to Redox Reactions

1.4.1 Oxidation and reduction definitions

Oxidation is the process in which a substance loses electrons, while **reduction** is the process in which a substance gains electrons. These two processes always occur together in a chemical

reaction, hence the term **redox reaction**. For example, in the reaction between zinc and copper(II) sulphate, zinc is oxidised as it loses electrons, and copper ions are reduced as they gain electrons.

Redox reactions are important because they explain processes such as respiration, combustion, and corrosion.

Fill in the blank: A reaction in which one substance loses electrons and another gains electrons is called a _____ reaction.

1.4.2 Identifying oxidising and reducing agents

An **oxidising agent** is a substance that causes oxidation by accepting electrons from another substance. A **reducing agent** is a substance that causes reduction by donating electrons. In the zinc and copper(II) sulphate reaction, copper ions act as the oxidising agent, while zinc acts as the reducing agent.

It is important to identify these agents because they help us understand the direction of electron flow in reactions.

Fill in the blank: A substance that donates electrons in a redox reaction is called a _____ agent.

1.4.3 Balancing simple redox equations

Balancing redox equations ensures that both the number of atoms and the charges are equal on both sides of the equation. This is achieved using the **ion-electron method** or the **oxidation number method**.

For example, consider the reaction between iron (II) ions and dichromate ions in acidic solution. The half-reactions are written separately, balanced for atoms and charges, then combined to give the overall balanced equation.

Balancing redox equations requires careful step-by-step work. Always begin by writing the half-reactions, then balance atoms other than hydrogen and oxygen, followed by oxygen using water, hydrogen using hydrogen ions, and charges using electrons.

Fill in the blank: The method commonly used to balance redox equations by writing half-reactions is called the _____ method.

Pilot questions 1.4

1. Define oxidation and reduction in terms of electron transfer.
2. In the reaction between zinc and copper(II) sulphate, which substance acts as the oxidising agent?
3. What is the role of a reducing agent in a redox reaction?
4. Name one common oxidising agent and one common reducing agent.
5. Which method is commonly used to balance redox equations?

Series Summary

This series has introduced you to the fundamental concepts that form the basis of chemistry. Its purpose was to help you appreciate how atoms and molecules act as the building blocks of matter, how chemical reactions occur, and how these processes can be represented and analysed. By working through the content, you have gained the essential foundation needed to progress into more advanced areas of chemistry.

We began by examining atoms and molecules, then moved on to chemical reactions and equations, where you learnt how to classify, write, and balance them. Following that, we explored stoichiometry and quantitative relationships, focusing on the mole concept and its applications in problem solving. Finally, we concluded with an introduction to redox reactions, identifying oxidation and reduction processes and practising how to balance simple redox equations. In line with the Series Learning Outcomes, you should now be able to define atoms and molecules, explain the distinction between elements and compounds, describe and balance chemical equations, apply the mole concept to quantitative problems, and identify oxidation and reduction processes in redox reactions.

Concept Check Questions Series 1

Objective questions

1. Which subatomic particles are found in the nucleus of an atom? (Linked to SLO 2.1)
2. The chemical symbol for sodium is derived from its Latin name. What is that name? (Linked to SLO 2.2)
3. A reaction in which a single compound breaks down into simpler substances is called a _____ reaction. (Linked to SLO 2.3)
4. At standard temperature and pressure, one mole of any gas occupies how many litres? (Linked to SLO 2.5)
5. In a redox reaction, the substance that donates electrons is called the _____ agent. (Linked to SLO 2.7)

Short-answer questions

6. Define a molecule and give one example. (Linked to SLO 2.1)
7. Explain why chemical equations must be balanced. (Linked to SLO 2.4)
8. State the value of Avogadro's constant. (Linked to SLO 2.5)
9. Identify the oxidising agent in the reaction between zinc and copper(II) sulphate. (Linked to SLO 2.7)

10. Describe the difference between a combination reaction and a displacement reaction. (Linked to SLO 2.3)

Long-answer questions

11. Analyse the importance of stoichiometry in chemical problem solving, using at least two practical examples. (Linked to SLO 2.6)

12. Demonstrate how to balance a simple redox equation using the ion-electron method, and evaluate why this method is effective. (Linked to SLO 2.7)

Answers to Concept Check Questions [Series 1]

Objective questions

1. Protons and neutrons.
2. *Natrium*.
3. Decomposition reaction.
4. 22.4 litres.
5. Reducing agent.

Short-answer questions

6. A molecule is a group of two or more atoms held together by chemical bonds. Example: O_2.

7. Chemical equations must be balanced to obey the law of conservation of mass, ensuring equal numbers of atoms on both sides.

8. 6.022×1023particles per mole.

9. Copper(II) ions act as the oxidising agent.

10. A combination reaction involves two or more substances forming one product, while a displacement reaction involves one element replacing another in a compound.

Long-answer questions

11. *Basic response*: Stoichiometry is important because it allows chemists to calculate the amounts of reactants and products in a reaction. For example, it helps determine the mass of oxygen needed to completely burn hydrogen, or the concentration of an acid in titration experiments. *Value-added response*: Beyond the classroom, stoichiometry is vital in industrial processes such as fertiliser production, where precise ratios of nitrogen compounds are required, and in environmental chemistry, where stoichiometric calculations help assess pollutant levels and design treatment methods.

12. *Basic response*: To balance a redox equation using the ion-electron method, write the half-reactions, balance atoms other than hydrogen and oxygen, balance oxygen with water, balance

hydrogen with H^+, balance charges with electrons, then combine the half-reactions. This ensures both mass and charge are conserved. *Value-added response*: The ion-electron method is effective because it systematically accounts for both matter and charge, making it applicable to complex reactions in acidic or basic solutions. It is widely used in electrochemistry, for example in balancing reactions in batteries and fuel cells, where accurate electron transfer calculations are essential for predicting efficiency.

Answers to Pilot Questions [Series 1]

Part 1: Atoms and Molecules: The Building Blocks of Matter

- Protons and neutrons.
- Atom.
- Oxygen (O_2).
- A molecule is made up of atoms bonded together, while a compound consists of atoms of different elements combined in fixed proportions.
- Cl.

Part 2: Chemical Reactions and Equations

- Reactants are substances that undergo change, while products are the new substances formed.
- $2H_2 + O_2 \rightarrow 2H_2O$.
- Exchange (double displacement) reaction.
- To obey the law of conservation of mass.
- $2H_2 + O_2 \rightarrow 2H_2O$.

Part 3: Stoichiometry and Quantitative Relationships

- 6.022×1023.
- 6.022×1023molecules.
- 22.4 litres.
- 32 g of oxygen.
- Correct proportions.

Part 4: Introduction to Redox Reactions

- Oxidation is loss of electrons; reduction is gain of electrons.
- Copper (II) ions.
- Reducing agent.
- Oxygen (oxidising agent), hydrogen (reducing agent).

- Ion-electron method.

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Course Unit: 2 Units

Series 2: Electronic Structure, Periodicity, and Molecular Shapes

Part 1 Modern electronic theory of atoms

- Structure of the atom and energy levels
- Quantum numbers and orbitals
- Aufbau principle, Pauli exclusion principle, and Hund's rule

Part 2 Electronic configurations of elements

- Writing electronic configurations for main group elements
- Transition elements and exceptions
- Applications of electronic configurations in chemical behaviour

Part 3 Periodicity and trends in the periodic table

- Atomic radii and ionic radii
- Ionisation energy and electron affinity
- Electronegativity and chemical reactivity

Part 4 Molecular shapes and hybridisation

- Valence Shell Electron Pair Repulsion (VSEPR) theory
- Common molecular geometries (linear, trigonal planar, tetrahedral, etc.)
- Hybridisation of orbitals and examples

Introduction

In the last Series, we explored the foundations of atoms, molecules, and chemical reactions, which provided you with the basic building blocks of matter and the processes by which substances interact. Building on that knowledge, this Series will take you deeper into the structure of atoms by focusing on how electrons are arranged, how these arrangements explain periodic trends, and how they influence the shapes of molecules. These concepts are central to understanding chemical behaviour and predicting the properties of elements and compounds.

The aim of this Series is to equip you with the ability to describe electronic structures, analyse periodic trends, and demonstrate how molecular shapes and hybridisation arise from electron arrangements. By mastering these topics, you will gain the tools needed to explain why elements behave as they do and how molecules adopt specific geometries.

We shall commence this Series by examining the modern electronic theory of atoms, focusing on energy levels, quantum numbers, and the principles that govern electron arrangements. Next, we will move on to writing electronic configurations for elements, including main group and transition elements, and consider how these configurations relate to chemical behaviour. Following that, we will explore periodicity and trends in the periodic table, such as atomic radii, ionisation energy, and electronegativity. Finally, we will wrap up by studying molecular shapes and hybridisation, using VSEPR theory to predict geometries and examining how orbitals combine to form hybridised structures.

Series Learning Outcomes

Upon completing this Series, you should be able to:

- **SLO 2.1** explain the modern electronic theory of atoms, including energy levels, quantum numbers, and orbitals,
- **SLO 2.2** demonstrate how to apply the Aufbau principle, Pauli exclusion principle, and Hund's rule in predicting electron arrangements,
- **SLO 2.3** write electronic configurations for main group and transition elements, including recognising exceptions,
- **SLO 2.4** analyse how electronic configurations influence the chemical behaviour of elements,
- **SLO 2.5** evaluate periodic trends such as atomic radii, ionisation energy, electron affinity, and electronegativity,
- **SLO 2.6** apply the Valence Shell Electron Pair Repulsion (VSEPR) theory to predict molecular geometries, and
- **SLO 2.7** describe the concept of hybridisation and illustrate how orbitals combine to form common molecular shapes.

2.1 Modern Electronic Theory of Atoms

2.1.1 Structure of the atom and energy levels

An **atom** is the smallest unit of matter that retains the properties of an element. It consists of a central **nucleus**, made up of positively charged protons and neutral neutrons, surrounded by negatively charged electrons that occupy specific regions called **energy levels** or shells. These

energy levels are quantised, meaning electrons can only exist at certain discrete energy states rather than in between.

The arrangement of electrons in these energy levels determines the chemical behaviour of the atom. For example, atoms with a full outer shell are generally stable, while those with incomplete shells tend to react to achieve stability.

Fill in the blank: Electrons in an atom occupy discrete regions called _____ levels.

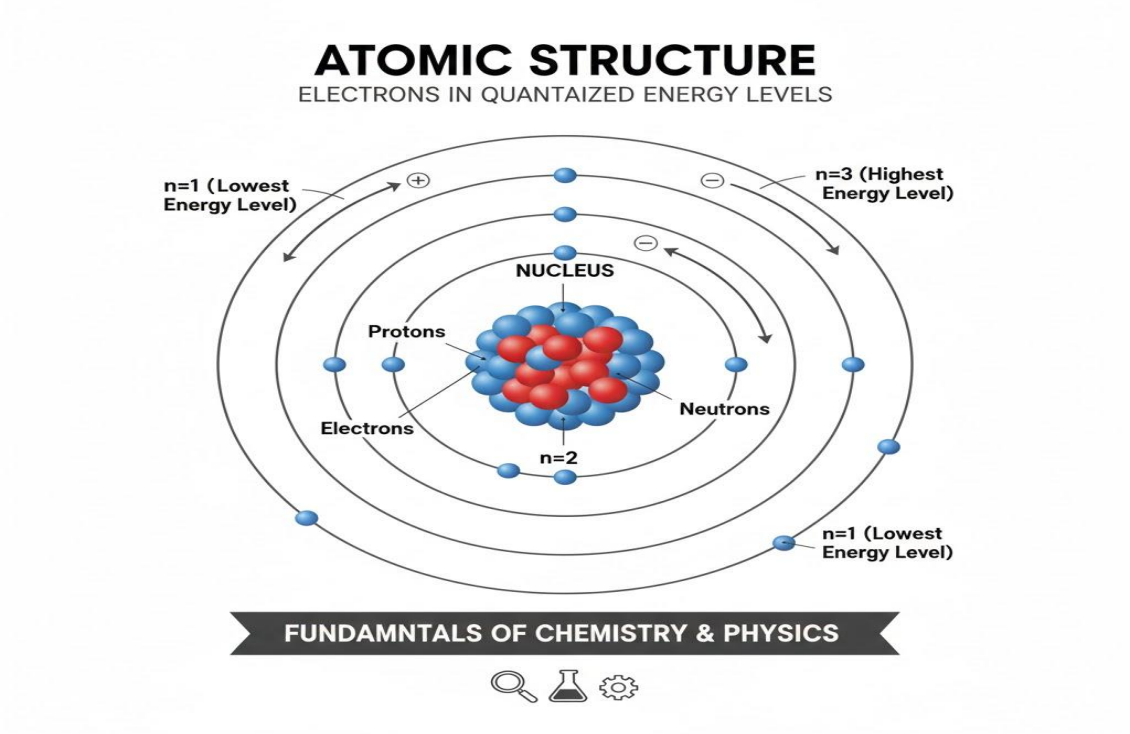


Figure 2.1: Structure of an atom with energy levels

2.1.2 Quantum numbers and orbitals

Electrons are further described using **quantum numbers**, which specify their position and behaviour within an atom. There are four quantum numbers:

- **Principal quantum number (n):** Indicates the main energy level or shell.
- **Azimuthal quantum number (l):** Defines the shape of the orbital (s, p, d, f).
- **Magnetic quantum number (m):** Specifies the orientation of the orbital in space.
- **Spin quantum number (s):** Describes the spin of the electron, either +½ or −½.

An **orbital** is a region in space where there is a high probability of finding an electron. Orbitals have different shapes depending on their type: s orbitals are spherical, p orbitals are dumbbell-shaped, and d orbitals have more complex forms.

Fill in the blank: The quantum number that defines the shape of an orbital is called the _____ quantum number.

2.1.3 Aufbau principle, Pauli exclusion principle, and Hund's rule

The arrangement of electrons in orbitals follows specific rules:

- **Aufbau principle:** Electrons fill orbitals starting from the lowest energy level to higher ones.
- **Pauli exclusion principle:** No two electrons in an atom can have the same set of four quantum numbers, meaning each orbital can hold a maximum of two electrons with opposite spins.
- **Hund's rule:** When filling orbitals of equal energy (degenerate orbitals), electrons occupy them singly first with parallel spins before pairing.

These principles ensure that electrons are distributed in the most stable arrangement possible. For example, the electronic configuration of oxygen (atomic number 8) is $1s^2\ 2s^2\ 2p^4$, following these rules.

Fill in the blank: According to Hund's rule, electrons occupy orbitals of equal energy _____ before pairing.

Pilot questions 2.1

1. What particles are found in the nucleus of an atom?
2. Name the four quantum numbers used to describe electrons.
3. What is the shape of a p orbital?
4. State the Aufbau principle.
5. According to Hund's rule, how do electrons fill orbitals of equal energy?

2.2 Electronic Configurations of Elements

2.2.1 Writing electronic configurations for main group elements

An **electronic configuration** is the arrangement of electrons in the orbitals of an atom. It shows how electrons are distributed among energy levels and sublevels (s, p, d, f). For main group elements, configurations are written by following the Aufbau principle, Pauli exclusion principle, and Hund's rule.

For example, the configuration of carbon (atomic number 6) is $1s^2\ 2s^2\ 2p^2$. This means two electrons occupy the 1s orbital, two in the 2s orbital, and two in the 2p orbital.

Fill in the blank: The arrangement of electrons in the orbitals of an atom is called its electronic _____.

Element	Atomic Number	Electronic Configuration
Hydrogen (H)	1	$1s^1$
Helium (He)	2	$1s^2$

Lithium (Li)	3	$1s^2\ 2s^1$
Beryllium (Be)	4	$1s^2\ 2s^2$
Boron (B)	5	$1s^2\ 2s^2\ 2p^1$
Carbon (C)	6	$1s^2\ 2s^2\ 2p^2$
Nitrogen (N)	7	$1s^2\ 2s^2\ 2p^3$
Oxygen (O)	8	$1s^2\ 2s^2\ 2p^4$
Fluorine (F)	9	$1s^2\ 2s^2\ 2p^5$
Neon (Ne)	10	$1s^2\ 2s^2\ 2p^6$
Sodium (Na)	11	$1s^2\ 2s^2\ 2p^6\ 3s^1$
Magnesium (Mg)	12	$1s^2\ 2s^2\ 2p^6\ 3s^2$
Aluminium (Al)	13	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
Silicon (Si)	14	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
Phosphorus (P)	15	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
Sulfur (S)	16	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
Chlorine (Cl)	17	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
Argon (Ar)	18	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Table 2.2: Examples of electronic configurations of main group elements

2.2.2 Transition elements and exceptions

Transition elements often show exceptions to the expected filling order of orbitals. This is because half-filled and fully filled d orbitals provide extra stability. For example, chromium (atomic number 24) has the configuration [Ar] $3d^5\ 4s^1$ instead of [Ar] $3d^4\ 4s^2$, and copper (atomic number 29) has [Ar] $3d^{10}\ 4s^1$ instead of [Ar] $3d^9\ 4s^2$.

These exceptions highlight the importance of stability in electron arrangements.

Fill in the blank: The electronic configuration of copper is [Ar] $3d^{10}$ _____.

2.2.3 Applications of electronic configurations in chemical behaviour

Electronic configurations explain many properties of elements, such as their valency, reactivity, and placement in the periodic table. For instance, alkali metals have one electron in their outermost s orbital, making them highly reactive. Noble gases have full outer shells, which makes them stable and chemically inert.

Understanding configurations also helps predict bonding patterns. For example, oxygen with six valence electrons tends to form two covalent bonds to achieve a stable octet.

Fill in the blank: Elements with full outer shells, such as neon, are chemically _____.

Pilot questions 2.2

1. Write the electronic configuration of carbon.
2. What is the electronic configuration of sodium?

3. Why does chromium have the configuration [Ar] $3d^5$ $4s^1$ instead of [Ar] $3d^4$ $4s^2$?
4. Which group of elements has full outer shells and is chemically inert?
5. How does the electronic configuration of oxygen explain its bonding behaviour?

2.3 Periodicity and Trends in the Periodic Table

2.3.1 Atomic radii and ionic radii

The **atomic radius** is the distance from the nucleus of an atom to the outermost electron. It provides a measure of the size of an atom. The **ionic radius** refers to the size of an ion, which changes depending on whether the atom loses or gains electrons. Cations (positively charged ions) are smaller than their parent atoms because they lose electrons, while anions (negatively charged ions) are larger because they gain electrons.

Across a period in the periodic table, atomic radii decrease due to increasing nuclear charge pulling electrons closer to the nucleus. Down a group, atomic radii increase because additional electron shells are added.

Fill in the blank: When an atom loses electrons to form a cation, its ionic radius becomes _____ than the atomic radius.

2.3.2 Ionisation energy and electron affinity

Ionisation energy is the amount of energy required to remove an electron from an atom in its gaseous state. It increases across a period due to stronger nuclear attraction and decreases down a group as electrons are farther from the nucleus.

Electron affinity is the energy change when an atom gains an electron. Elements with high electron affinity, such as halogens, readily accept electrons to form negative ions.

Fill in the blank: The energy required to remove an electron from an atom is called _____ energy.

2.3.3 Electronegativity and chemical reactivity

Electronegativity is the ability of an atom to attract electrons in a chemical bond. It increases across a period and decreases down a group. Fluorine is the most electronegative element.

Periodic trends in electronegativity explain chemical reactivity. For example, alkali metals are highly reactive because of their low electronegativity and tendency to lose electrons, while halogens are reactive due to their high electronegativity and tendency to gain electrons.

Fill in the blank: The most electronegative element in the periodic table is _____.

Pilot questions 2.3

1. Define atomic radius and explain how it changes across a period.
2. Why are cations smaller than their parent atoms?
3. What is ionisation energy?
4. Which group of elements has the highest electron affinity?
5. Name the most electronegative element in the periodic table.

2.4 Molecular Shapes and Hybridisation

2.4.1 Valence Shell Electron Pair Repulsion (VSEPR) theory

The **Valence Shell Electron Pair Repulsion (VSEPR) theory** explains the shapes of molecules based on the repulsion between electron pairs around a central atom. Electron pairs, whether bonding or non-bonding, arrange themselves as far apart as possible to minimise repulsion. This arrangement determines the geometry of the molecule.

For example, methane (CH_4) has four bonding pairs around carbon, which arrange themselves in a tetrahedral geometry. Water (H_2O), with two bonding pairs and two lone pairs, adopts a bent shape.

Fill in the blank: According to VSEPR theory, electron pairs around a central atom arrange themselves to minimise _____.

2.4.2 Common molecular geometries

Molecules adopt different geometries depending on the number of bonding and non-bonding pairs around the central atom. Some common geometries include:

- **Linear:** Two bonding pairs, 180° apart (e.g., CO_2).
- **Trigonal planar:** Three bonding pairs, 120° apart (e.g., BF_3).
- **Tetrahedral:** Four bonding pairs, 109.5° apart (e.g., CH_4).
- **Trigonal bipyramidal:** Five bonding pairs, with 90° and 120° angles (e.g., PCl_5).
- **Octahedral:** Six bonding pairs, 90° apart (e.g., SF_6).

Fill in the blank: A molecule with four bonding pairs around the central atom adopts a _____ geometry.

2.4.3 Hybridisation of orbitals and examples

Hybridisation is the process by which atomic orbitals mix to form new hybrid orbitals that are suitable for bonding. This concept explains molecular shapes and bond strengths.

- **sp hybridisation:** One s orbital mixes with one p orbital, forming two linear orbitals (e.g., $BeCl_2$).

- **sp^2 hybridisation:** One s orbital mixes with two p orbitals, forming three trigonal planar orbitals (e.g., BF_3).
- **sp^3 hybridisation:** One s orbital mixes with three p orbitals, forming four tetrahedral orbitals (e.g., CH_4).
- **sp^3d hybridisation:** One s, three p, and one d orbital mix to form trigonal bipyramidal orbitals (e.g., PCl_5).
- **sp^3d^2 hybridisation:** One s, three p, and two d orbitals mix to form octahedral orbitals (e.g., SF_6).

Fill in the blank: The hybridisation that explains the tetrahedral geometry of methane is _____.

Pilot questions 2.4

1. State the main idea of VSEPR theory.
2. What is the molecular geometry of carbon dioxide?
3. Give one example of a molecule with trigonal planar geometry.
4. Which geometry has bond angles of 109.5°?
5. What type of hybridisation explains the tetrahedral shape of methane?

Series Summary

This Series has focused on how the arrangement of electrons within atoms governs their properties, explains periodic trends, and determines the shapes of molecules. Its purpose was to build on the foundations of atomic theory and chemical reactions by showing how electronic structure connects directly to chemical behaviour and molecular geometry, making it a vital step in your study of chemistry.

We began with the modern electronic theory of atoms, exploring energy levels, quantum numbers, and the principles that guide electron arrangements. Then we examined how to write electronic configurations for both main group and transition elements, including notable exceptions, and considered their applications in predicting chemical behaviour. Following that, we studied periodicity, looking at trends such as atomic radii, ionisation energy, electron affinity, and electronegativity. Finally, we concluded with molecular shapes and hybridisation, using VSEPR theory to predict geometries and orbital mixing to explain bonding. In line with the Series Learning Outcomes, you should now be able to explain atomic structure, demonstrate electron configurations, analyse periodic trends, and apply theories of molecular geometry and hybridisation to chemical systems.

Concept Check Questions [Series 2]

Objective questions

1. The quantum number that defines the shape of an orbital is the _____ quantum number. (Linked to SLO 2.1)
2. Which principle states that electrons fill orbitals starting from the lowest energy level? (Linked to SLO 2.2)
3. The electronic configuration of sodium is _____. (Linked to SLO 2.3)
4. Across a period, atomic radius generally _____ due to increasing nuclear charge. (Linked to SLO 2.5)
5. The most electronegative element in the periodic table is _____. (Linked to SLO 2.5)

Short-answer questions

6. Define an orbital and give one example. (Linked to SLO 2.1)
7. State Hund's rule and explain its significance. (Linked to SLO 2.2)
8. Write the electronic configuration of oxygen. (Linked to SLO 2.3)
9. Explain why cations are smaller than their parent atoms. (Linked to SLO 2.5)
10. What is the molecular geometry of methane according to VSEPR theory? (Linked to SLO 2.6)

Long-answer questions

11. Analyse how electronic configurations influence the chemical behaviour of elements, using examples from alkali metals and noble gases. (Linked to SLO 2.4)
12. Demonstrate how VSEPR theory predicts molecular shapes, and evaluate its usefulness by comparing methane and water. (Linked to SLO 2.6)
13. Describe the concept of hybridisation and illustrate how it explains the bonding in methane and ethene. (Linked to SLO 2.7)

Answers to Concept Check Questions [Series 2]

Objective questions

1. Azimuthal quantum number.
2. Aufbau principle.
3. $1s^2\ 2s^2\ 2p^6\ 3s^1$.
4. Decreases.
5. Fluorine.

Short-answer questions

6. An orbital is a region in space where there is a high probability of finding an electron. Example: 2p orbital.

7. Hund's rule states that electrons occupy orbitals of equal energy singly with parallel spins before pairing. This minimises repulsion and stabilises the atom.

8. $1s^2\ 2s^2\ 2p^4$.

9. Cations are smaller because they lose electrons, reducing electron repulsion and allowing the nucleus to pull remaining electrons closer.

10. Tetrahedral geometry.

Long-answer questions

11. *Basic response*: Electronic configurations influence chemical behaviour by determining valency and reactivity. Alkali metals with one outer electron are highly reactive, while noble gases with full outer shells are inert. *Value-added response*: Beyond basic reactivity, electronic configurations explain trends in ionisation energy and bonding. For example, alkali metals readily form ionic compounds due to low ionisation energy, while noble gases are used in applications like lighting because of their stability.

12. *Basic response*: VSEPR theory predicts shapes by minimising electron pair repulsion. Methane has four bonding pairs around carbon, giving a tetrahedral shape, while water has two bonding pairs and two lone pairs, resulting in a bent shape. *Value-added response*: VSEPR theory is useful for predicting simple geometries but has limitations in complex molecules. For instance, it does not fully account for orbital hybridisation, which provides a deeper explanation of bond angles and strengths.

13. *Basic response*: Hybridisation is the mixing of atomic orbitals to form new orbitals for bonding. Methane involves sp^3 hybridisation, giving a tetrahedral shape, while ethene involves sp^2 hybridisation, giving a trigonal planar shape. *Value-added response*: Hybridisation explains not only shapes but also bond strengths and overlaps. In ethene, sp^2 hybridisation allows for strong sigma bonds and a pi bond, accounting for its reactivity in addition reactions.

Answers to Pilot Questions [Series 2]

Part 1: Modern Electronic Theory of Atoms

- Protons and neutrons.
- Principal, azimuthal, magnetic, and spin quantum numbers.
- Dumbbell-shaped.
- Electrons fill orbitals starting from the lowest energy level.
- Singly.

Part 2: Electronic Configurations of Elements

- $1s^2\ 2s^2\ 2p^2$.

- $1s^2\ 2s^2\ 2p^6\ 3s^1$.
- Because half-filled d orbitals provide extra stability.
- Noble gases.
- Oxygen tends to form two covalent bonds to complete its octet.

Part 3: Periodicity and Trends in the Periodic Table

- Atomic radius is the distance from the nucleus to the outermost electron; it decreases across a period.
- Because they lose electrons, reducing electron repulsion.
- The energy required to remove an electron from an atom in its gaseous state.
- Halogens.
- Fluorine.

Part 4: Molecular Shapes and Hybridisation

- Electron pairs arrange themselves to minimise repulsion.
- Linear geometry.
- BF_3.
- Tetrahedral geometry.
- sp^3 hybridisation.

Course code: CHM 101

Course Title: General Chemistry I

Course Unit: 2 Units

Series 3: Chemical Bonding, States of Matter, and Thermochemistry

Part 1 Chemical bonding

- Types of chemical bonds (ionic, covalent, metallic)
- Bond polarity and electronegativity
- Bond energy and bond length
- Intermolecular forces (hydrogen bonding, van der Waals forces)

Part 2 States of matter

- Properties of solids, liquids, and gases
- Gas laws (Boyle's, Charles's, Avogadro's, Ideal Gas Law)

- Real gases and deviations from ideal behaviour
- Phase changes and phase diagrams

Part 3 Thermochemistry

- Heat, work, and internal energy
- Enthalpy and calorimetry
- Hess's law and standard enthalpies of formation
- Bond enthalpies and energy changes in reactions

Introduction

In the last Series, we examined how electrons are arranged in atoms, how periodic trends arise, and how molecular shapes can be predicted. That knowledge provides the foundation for understanding how atoms interact with one another to form compounds, how matter exists in different physical states, and how energy changes accompany chemical processes. This Series will therefore connect microscopic bonding interactions with macroscopic properties and energetic considerations, making it central to your grasp of general chemistry.

The aim of this Series is to introduce you to the principles of chemical bonding, the behaviour of matter in its various states, and the fundamentals of thermochemistry. By the end, you should be able to explain how bonds form, describe the properties of solids, liquids, and gases, and analyse energy changes in chemical reactions.

We shall commence this Series by exploring chemical bonding, focusing on ionic, covalent, and metallic bonds, as well as bond polarity, bond energy, and intermolecular forces. Next, we will move on to the states of matter, considering the properties of solids, liquids, and gases, the gas laws, real gas behaviour, and phase changes. Finally, we will wrap up with thermochemistry, where we will study heat, work, enthalpy, calorimetry, Hess's law, and bond enthalpies, bringing the Series to a close with a clear understanding of how energy is involved in chemical reactions.

Series Learning Outcomes

Upon completing this Series, you should be able to:

- **SLO 3.1**: Define the different types of chemical bonds, including ionic, covalent, and metallic bonds.
- **SLO 3.2**: Explain the concepts of bond polarity, bond energy, bond length, and intermolecular forces.

- **SLO 3.3**: Describe the properties of solids, liquids, and gases, and distinguish between them.
- **SLO 3.4**: Apply the gas laws to calculate relationships among pressure, volume, temperature, and amount of gas.
- **SLO 3.5**: Analyse deviations of real gases from ideal behaviour using appropriate models.
- **SLO 3.6**: Demonstrate how to interpret phase changes and phase diagrams.
- **SLO 3.7**: Define heat, work, and internal energy in thermochemical processes.
- **SLO 3.8**: Calculate enthalpy changes using calorimetry and Hess's law.
- **SLO 3.9**: Evaluate energy changes in chemical reactions using bond enthalpies.

3.1 Chemical Bonding

3.1.1 Types of chemical bonds

A **chemical bond** is the force of attraction that holds atoms or ions together in a compound. The three main types are:

- **Ionic bonds:** These occur when electrons are transferred from one atom to another, usually between metals and non-metals. The resulting cations and anions are held together by electrostatic attraction.
- **Covalent bonds:** These involve the sharing of electron pairs between atoms, typically between non-metals. The shared electrons allow each atom to achieve a stable configuration.
- **Metallic bonds:** Found in metals, these bonds involve delocalised valence electrons moving freely through a lattice of positive ions, giving rise to properties such as conductivity and malleability.

Fill in the blank: A bond formed by the sharing of electron pairs between atoms is called a _____ bond.

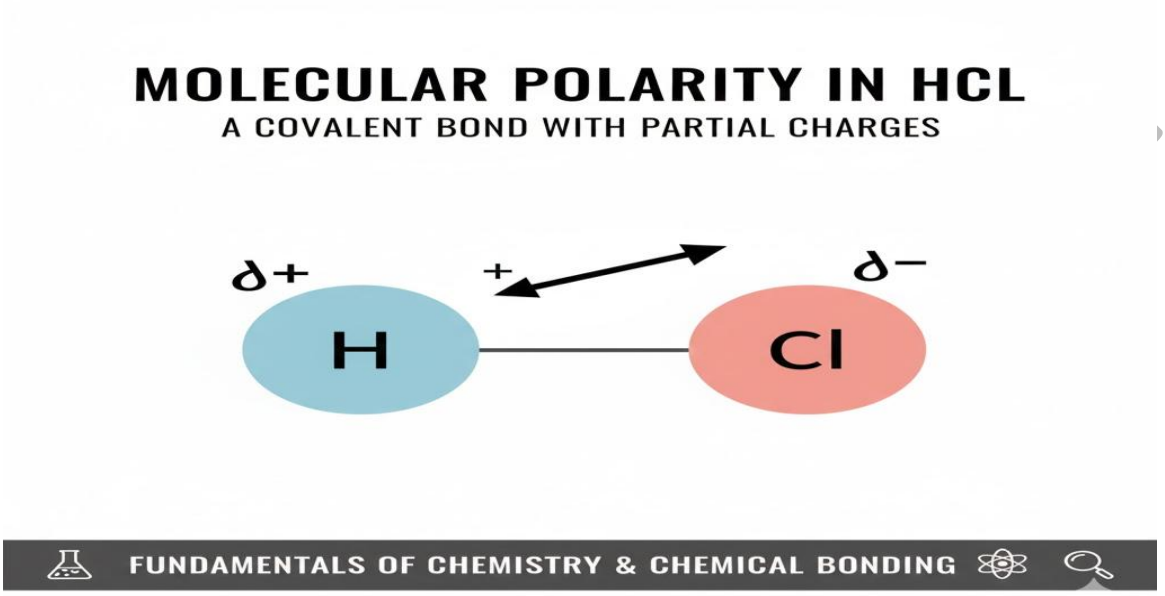
3.1.2 Bond polarity and electronegativity

Electronegativity is the ability of an atom to attract electrons in a chemical bond. When two atoms with different electronegativities form a covalent bond, the electrons are shared unequally, creating a **polar bond**. If the difference in electronegativity is very large, the bond may become ionic.

For example, in hydrogen chloride (HCl), chlorine is more electronegative than hydrogen, so the bond is polar with partial charges on each atom.

Fill in the blank: When electrons are shared unequally between atoms, the bond is described as _____

Plate 3.1: Polarity in a covalent bond



3.1.3 Bond energy and bond length

Bond energy is the amount of energy required to break one mole of a bond in the gaseous state. It reflects the strength of the bond. **Bond length** is the average distance between the nuclei of two bonded atoms. Stronger bonds generally have shorter bond lengths.

For example, a carbon–carbon triple bond (C≡C) has higher bond energy and shorter bond length than a carbon–carbon double bond (C=C).

Fill in the blank: The average distance between the nuclei of two bonded atoms is called bond _____.

Bond Type	Bond Energy (kJ/mol)	Bond Length (nm)
C–C	331	0.154
C=C	590	0.134
C≡C	812	0.120

Table 3.1: Bond energies and bond lengths of selected bonds

3.1.4 Intermolecular forces

Intermolecular forces are weak forces of attraction between molecules. They are not chemical bonds but strongly influence physical properties such as boiling and melting points.

- **Van der Waals forces:** Weak attractions due to temporary dipoles.
- **Dipole–dipole interactions:** Attractions between polar molecules.
- **Hydrogen bonding:** A strong type of dipole–dipole interaction involving hydrogen bonded to highly electronegative atoms such as oxygen, nitrogen, or fluorine.

For example, hydrogen bonding in water explains its unusually high boiling point compared to other molecules of similar size.

Fill in the blank: The unusually high boiling point of water is due to _____ bonding.

Pilot questions 3.1

1. What type of bond is formed when electrons are transferred between atoms?
2. Define electronegativity.
3. What is the difference between bond energy and bond length?
4. Name the three main types of intermolecular forces.
5. Why does water have a high boiling point compared to similar molecules?

3.2 States of Matter

3.2.1 Properties of solids, liquids, and gases

Matter exists in three main physical states: **solids, liquids, and gases**.

- **Solids:** Particles are closely packed in fixed positions, giving solids a definite shape and volume. They are rigid and incompressible.
- **Liquids:** Particles are close together but can move past one another, so liquids have a definite volume but take the shape of their container. They are less rigid and slightly compressible.
- **Gases:** Particles are far apart and move freely, so gases have neither definite shape nor volume. They are highly compressible and expand to fill their container.

Fill in the blank: A state of matter with definite volume but no fixed shape is called a _____.

3.2.2 Gas laws

The behaviour of gases can be described using several **gas laws**:

- **Boyle's law:** At constant temperature, the pressure of a gas is inversely proportional to its volume.
- **Charles's law:** At constant pressure, the volume of a gas is directly proportional to its temperature.
- **Avogadro's law:** Equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.
- **Ideal Gas Law:** Combines these relationships into the equation PV=nRT, where Pis pressure, Vis volume, nis moles of gas, Ris the gas constant, and Tis temperature.

Fill in the blank: The equation PV=nRTis known as the _____ Gas Law.

3.2.3 Real gases and deviations from ideal behaviour

The **Ideal Gas Law** assumes that gas particles have no volume and no intermolecular forces. However, real gases deviate from this behaviour, especially at high pressures and low temperatures.

The **van der Waals equation** modifies the Ideal Gas Law to account for particle volume and intermolecular forces:

$$(P+an2V2)(V-nb)=nRT$$

where aand bare constants specific to each gas.

Fill in the blank: Real gases deviate most from ideal behaviour at _____ pressure and _____ temperature.

3.2.4 Phase changes and phase diagrams

Phase changes are transitions between states of matter, such as melting, freezing, vaporisation, condensation, sublimation, and deposition. These processes involve energy changes but not chemical changes.

A **phase diagram** shows the conditions of temperature and pressure under which different states of matter exist. It includes the **triple point**, where solid, liquid, and gas coexist, and the **critical point**, beyond which a gas cannot be liquefied.

Fill in the blank: The point on a phase diagram where solid, liquid, and gas coexist is called the _____ point.

Pilot questions 3.2

1. What state of matter has definite shape and volume?
2. State Boyle's law in words.
3. Write the equation of the Ideal Gas Law.
4. Under what conditions do real gases deviate most from ideal behaviour?
5. What is the triple point in a phase diagram?

3.3 Thermochemistry

3.3.1 Heat, work, and internal energy

Thermochemistry is the study of energy changes, particularly heat, during chemical reactions. **Heat** is the transfer of energy due to a temperature difference, while **work** is energy transfer that results from a force acting through a distance. The **internal energy** of a system is the sum of all kinetic and potential energies of its particles.

In chemical reactions, energy can be absorbed or released. An endothermic process absorbs heat, while an exothermic process releases heat.

Fill in the blank: A chemical reaction that absorbs heat from its surroundings is called an _____ process.

3.3.2 Enthalpy and calorimetry

Enthalpy (H) is the heat content of a system at constant pressure. The change in enthalpy (H) indicates whether a reaction is endothermic ($H>0$) or exothermic ($H<0$).

Calorimetry is the experimental measurement of heat changes. A calorimeter is used to measure the heat absorbed or released during a chemical or physical process. For example, burning a fuel in a calorimeter allows us to determine its enthalpy of combustion.

Fill in the blank: The instrument used to measure heat changes in chemical reactions is called a _____.

3.3.3 Hess's law and standard enthalpies of formation

Hess's law states that the enthalpy change of a reaction is the same whether it occurs in one step or several steps. This principle allows us to calculate enthalpy changes indirectly.

Standard enthalpy of formation (Hf**)** is the enthalpy change when one mole of a compound is formed from its elements in their standard states. These values are tabulated and widely used in calculations.

Fill in the blank: The enthalpy change of a reaction is independent of the path taken, according to _____ law.

3.3.4 Bond enthalpies and energy changes in reactions

Bond enthalpy is the energy required to break one mole of a particular bond in the gaseous state. It provides a way to estimate the enthalpy change of a reaction by comparing the energy required to break bonds in reactants with the energy released when new bonds form in products.

For example, in the combustion of methane, C–H and O=O bonds are broken, while C=O and O–H bonds are formed. The difference in bond enthalpies gives the overall enthalpy change.

Fill in the blank: The energy required to break one mole of a bond in the gaseous state is called bond _____.

Pilot questions 3.3

1. Define internal energy.
2. What is the difference between endothermic and exothermic processes?
3. What instrument is used to measure heat changes in reactions?
4. State Hess's law.
5. What is bond enthalpy?

Series Summary

This series has explored how atoms bond together, how matter behaves in different physical states, and how energy changes accompany chemical reactions. Its purpose was to connect microscopic interactions with macroscopic properties and energetic principles, helping you see the direct link between bonding, states of matter, and thermochemistry in everyday chemical systems.

We began with chemical bonding, examining ionic, covalent, and metallic bonds, as well as bond polarity, bond energy, and intermolecular forces. We then moved on to the states of matter, considering the properties of solids, liquids, and gases, the gas laws, real gas behaviour, and phase diagrams. Finally, we concluded with thermochemistry, where we studied heat, work,

enthalpy, calorimetry, Hess's law, and bond enthalpies. In line with the Series Learning Outcomes, you should now be able to define and explain different types of bonds, apply gas laws, analyse deviations in real gases, interpret phase changes, calculate enthalpy changes, and evaluate energy transformations in chemical reactions.

Concept Check Questions [Series 3]

Objective questions

1. A bond formed by the transfer of electrons between atoms is called an _____ bond. (Linked to SLO 3.1)
2. The ability of an atom to attract electrons in a chemical bond is termed _____. (Linked to SLO 3.2)
3. At constant temperature, the pressure of a gas is inversely proportional to its volume. This statement describes _____ law. (Linked to SLO 3.4)
4. The point on a phase diagram where solid, liquid, and gas coexist is called the _____ point. (Linked to SLO 3.6)
5. The enthalpy change of a reaction is independent of the path taken, according to _____ law. (Linked to SLO 3.8)

Short-answer questions

6. Define metallic bonding and explain one property it accounts for in metals. (Linked to SLO 3.1)
7. Distinguish between bond energy and bond length. (Linked to SLO 3.2)
8. State Charles's law in words. (Linked to SLO 3.4)
9. Explain why real gases deviate from ideal behaviour at high pressure. (Linked to SLO 3.5)
10. What instrument is used to measure heat changes in chemical reactions? (Linked to SLO 3.7)

Long-answer questions

11. Analyse the role of intermolecular forces in determining physical properties, using water and carbon dioxide as examples. (Linked to SLO 3.2)
12. Demonstrate how to apply the Ideal Gas Law to calculate the volume of a gas at given conditions, and evaluate its limitations. (Linked to SLO 3.4, 3.5)
13. Describe Hess's law and illustrate its application with a worked example. (Linked to SLO 3.8)
14. Evaluate how bond enthalpies can be used to estimate reaction enthalpies, with reference to combustion reactions. (Linked to SLO 3.9)

Answers to Concept Check Questions [Series 3]

Objective questions

1. Ionic bond.
2. Electronegativity.
3. Boyle's law.
4. Triple point.
5. Hess's law.

Short-answer questions

6. Metallic bonding involves delocalised electrons moving freely among positive ions. It explains conductivity in metals.
7. Bond energy is the energy required to break one mole of a bond, while bond length is the average distance between nuclei of bonded atoms.
8. Charles's law states that at constant pressure, the volume of a gas is directly proportional to its temperature.
9. At high pressure, gas particles occupy significant volume and intermolecular forces become important, causing deviations.
10. Calorimeter.

Long-answer questions

11. *Basic response*: Intermolecular forces affect boiling and melting points. Water has hydrogen bonding, giving it a high boiling point, while carbon dioxide has weak van der Waals forces, making it a gas at room temperature. *Value-added response*: Intermolecular forces also influence solubility and biological processes. For example, hydrogen bonding stabilises DNA base pairs, while weak forces determine the behaviour of gases in the atmosphere.
12. *Basic response*: The Ideal Gas Law PV=nRTcan be used to calculate gas volume when pressure, temperature, and moles are known. *Value-added response*: While useful, the law assumes no intermolecular forces and negligible particle volume. At high pressures or low temperatures, real gases deviate, requiring corrections such as the van der Waals equation.
13. *Basic response*: Hess's law states that the enthalpy change of a reaction is the same regardless of the path taken. *Value-added response*: For example, the enthalpy of formation of CO_2 can be calculated indirectly by combining enthalpy changes of combustion reactions. This principle is widely applied in thermochemical cycles and industrial energy calculations.
14. *Basic response*: Bond enthalpies allow estimation of reaction enthalpies by comparing energy required to break bonds with energy released when new bonds form. *Value-added response*: In combustion reactions, such as burning methane, bond enthalpy calculations provide approximate values for energy released. These estimates are crucial in fuel efficiency studies and environmental impact assessments.

Answers to Pilot Questions [Series 3]

Part 3.1: Chemical Bonding

- Ionic bond.

- Electronegativity is the ability of an atom to attract electrons in a chemical bond.
- Bond energy is the energy required to break one mole of a bond, while bond length is the average distance between the nuclei of bonded atoms.
- Van der Waals forces, dipole–dipole interactions, and hydrogen bonding.
- Hydrogen bonding.

Part 3.2: States of Matter

- Solid.
- Boyle's law states that at constant temperature, the pressure of a gas is inversely proportional to its volume.
- PV=nRT.
- High pressure and low temperature.
- Triple point.

Part 3.3: Thermochemistry

- Internal energy is the sum of all kinetic and potential energies of the particles in a system.
- Endothermic processes absorb heat, while exothermic processes release heat.
- Calorimeter.
- Hess's law states that the enthalpy change of a reaction is the same regardless of the path taken.
- Bond enthalpy is the energy required to break one mole of a bond in the gaseous state.

Course code: CHM 101

Course Title: General Chemistry I

Course Unit: 2 Units

Series 2: Electronic Structure, Periodicity, and Molecular Shapes

Part 1 Modern electronic theory of atoms

- Structure of the atom and energy levels
- Quantum numbers and orbitals
- Aufbau principle, Pauli exclusion principle, and Hund's rule

Part 2 Electronic configurations of elements

- Writing electronic configurations for main group elements
- Transition elements and exceptions
- Applications of electronic configurations in chemical behaviour

Part 3 Periodicity and trends in the periodic table

- Atomic radii and ionic radii
- Ionisation energy and electron affinity
- Electronegativity and chemical reactivity

Part 4 Molecular shapes and hybridisation

- Valence Shell Electron Pair Repulsion (VSEPR) theory
- Common molecular geometries (linear, trigonal planar, tetrahedral, etc.)
- Hybridisation of orbitals and examples

Introduction

In the last Series, we explored the foundations of atoms, molecules, and chemical reactions, which provided you with the basic building blocks of matter and the processes by which substances interact. Building on that knowledge, this Series will take you deeper into the structure of atoms by focusing on how electrons are arranged, how these arrangements explain periodic trends, and how they influence the shapes of molecules. These concepts are central to understanding chemical behaviour and predicting the properties of elements and compounds.

The aim of this Series is to equip you with the ability to describe electronic structures, analyse periodic trends, and demonstrate how molecular shapes and hybridisation arise from electron arrangements. By mastering these topics, you will gain the tools needed to explain why elements behave as they do and how molecules adopt specific geometries.

We shall commence this Series by examining the modern electronic theory of atoms, focusing on energy levels, quantum numbers, and the principles that govern electron arrangements. Next, we will move on to writing electronic configurations for elements, including main group and transition elements, and consider how these configurations relate to chemical behaviour. Following that, we will explore periodicity and trends in the periodic table, such as atomic radii, ionisation energy, and electronegativity. Finally, we will wrap up by studying molecular shapes and hybridisation, using VSEPR theory to predict geometries and examining how orbitals combine to form hybridised structures.

Series Learning Outcomes

Upon completing this Series, you should be able to:

- **SLO 2.1** explain the modern electronic theory of atoms, including energy levels, quantum numbers, and orbitals,
- **SLO 2.2** demonstrate how to apply the Aufbau principle, Pauli exclusion principle, and Hund's rule in predicting electron arrangements,

- **SLO 2.3** write electronic configurations for main group and transition elements, including recognising exceptions,
- **SLO 2.4** analyse how electronic configurations influence the chemical behaviour of elements,
- **SLO 2.5** evaluate periodic trends such as atomic radii, ionisation energy, electron affinity, and electronegativity,
- **SLO 2.6** apply the Valence Shell Electron Pair Repulsion (VSEPR) theory to predict molecular geometries, and
- **SLO 2.7** describe the concept of hybridisation and illustrate how orbitals combine to form common molecular shapes.

2.1 Modern Electronic Theory of Atoms

2.1.1 Structure of the atom and energy levels

An **atom** is the smallest unit of matter that retains the properties of an element. It consists of a central **nucleus**, made up of positively charged protons and neutral neutrons, surrounded by negatively charged electrons that occupy specific regions called **energy levels** or shells. These energy levels are quantised, meaning electrons can only exist at certain discrete energy states rather than in between.

The arrangement of electrons in these energy levels determines the chemical behaviour of the atom. For example, atoms with a full outer shell are generally stable, while those with incomplete shells tend to react to achieve stability.

Fill in the blank: Electrons in an atom occupy discrete regions called _____ levels.

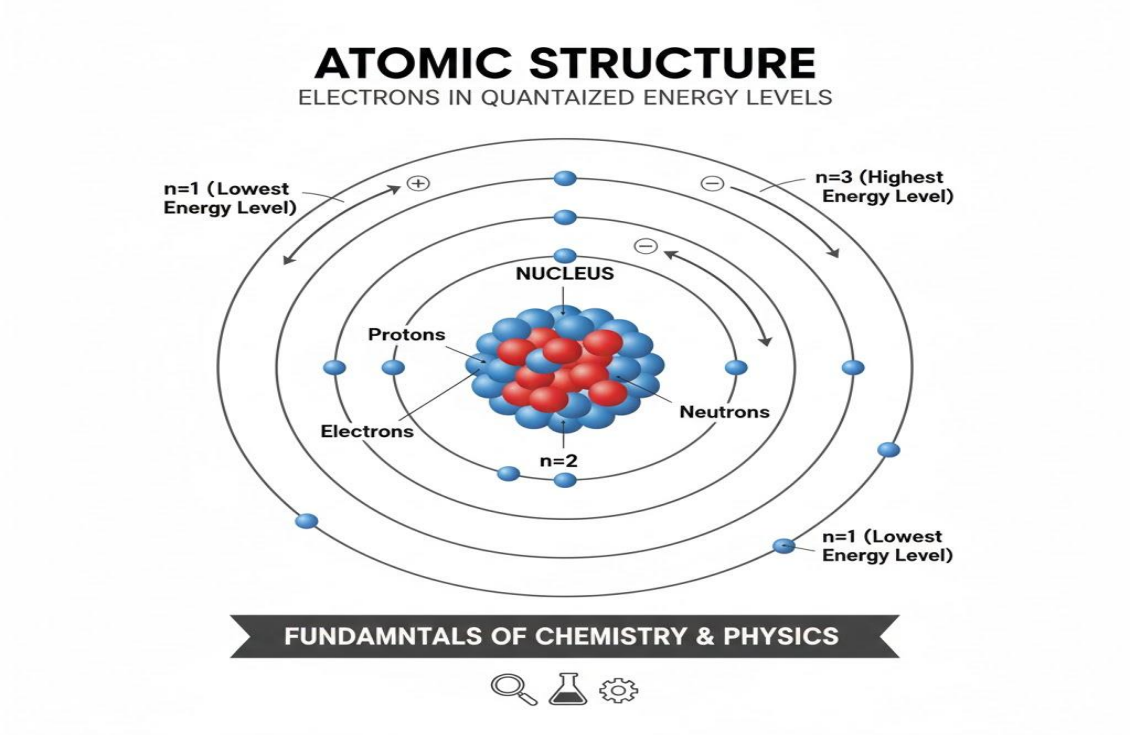


Figure 2.1: Structure of an atom with energy levels

2.1.2 Quantum numbers and orbitals

Electrons are further described using **quantum numbers**, which specify their position and behaviour within an atom. There are four quantum numbers:

- **Principal quantum number (n):** Indicates the main energy level or shell.
- **Azimuthal quantum number (l):** Defines the shape of the orbital (s, p, d, f).
- **Magnetic quantum number (m):** Specifies the orientation of the orbital in space.
- **Spin quantum number (s):** Describes the spin of the electron, either +½ or −½.

An **orbital** is a region in space where there is a high probability of finding an electron. Orbitals have different shapes depending on their type: s orbitals are spherical, p orbitals are dumbbell-shaped, and d orbitals have more complex forms.

Fill in the blank: The quantum number that defines the shape of an orbital is called the _____ quantum number.

2.1.3 Aufbau principle, Pauli exclusion principle, and Hund's rule

The arrangement of electrons in orbitals follows specific rules:

- **Aufbau principle:** Electrons fill orbitals starting from the lowest energy level to higher ones.
- **Pauli exclusion principle:** No two electrons in an atom can have the same set of four quantum numbers, meaning each orbital can hold a maximum of two electrons with opposite spins.
- **Hund's rule:** When filling orbitals of equal energy (degenerate orbitals), electrons occupy them singly first with parallel spins before pairing.

These principles ensure that electrons are distributed in the most stable arrangement possible. For example, the electronic configuration of oxygen (atomic number 8) is $1s^2\ 2s^2\ 2p^4$, following these rules.

Fill in the blank: According to Hund's rule, electrons occupy orbitals of equal energy _____ before pairing.

Pilot questions 2.1

1. What particles are found in the nucleus of an atom?
2. Name the four quantum numbers used to describe electrons.
3. What is the shape of a p orbital?
4. State the Aufbau principle.
5. According to Hund's rule, how do electrons fill orbitals of equal energy?

2.2 Electronic Configurations of Elements

2.2.1 Writing electronic configurations for main group elements

An **electronic configuration** is the arrangement of electrons in the orbitals of an atom. It shows how electrons are distributed among energy levels and sublevels (s, p, d, f). For main group elements, configurations are written by following the Aufbau principle, Pauli exclusion principle, and Hund's rule.

For example, the configuration of carbon (atomic number 6) is $1s^2\ 2s^2\ 2p^2$. This means two electrons occupy the 1s orbital, two in the 2s orbital, and two in the 2p orbital.

Fill in the blank: The arrangement of electrons in the orbitals of an atom is called its electronic _____.

Element	Atomic Number	Electronic Configuration
Hydrogen (H)	1	$1s^1$
Helium (He)	2	$1s^2$
Lithium (Li)	3	$1s^2\ 2s^1$
Beryllium (Be)	4	$1s^2\ 2s^2$
Boron (B)	5	$1s^2\ 2s^2\ 2p^1$
Carbon (C)	6	$1s^2\ 2s^2\ 2p^2$

Nitrogen (N)	7	$1s^2\ 2s^2\ 2p^3$
Oxygen (O)	8	$1s^2\ 2s^2\ 2p^4$
Fluorine (F)	9	$1s^2\ 2s^2\ 2p^5$
Neon (Ne)	10	$1s^2\ 2s^2\ 2p^6$
Sodium (Na)	11	$1s^2\ 2s^2\ 2p^6\ 3s^1$
Magnesium (Mg)	12	$1s^2\ 2s^2\ 2p^6\ 3s^2$
Aluminium (Al)	13	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
Silicon (Si)	14	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
Phosphorus (P)	15	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
Sulfur (S)	16	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
Chlorine (Cl)	17	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
Argon (Ar)	18	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Table 2.2: Examples of electronic configurations of main group elements

2.2.2 Transition elements and exceptions

Transition elements often show exceptions to the expected filling order of orbitals. This is because half-filled and fully filled d orbitals provide extra stability. For example, chromium (atomic number 24) has the configuration [Ar] $3d^5\ 4s^1$ instead of [Ar] $3d^4\ 4s^2$, and copper (atomic number 29) has [Ar] $3d^{10}\ 4s^1$ instead of [Ar] $3d^9\ 4s^2$.

These exceptions highlight the importance of stability in electron arrangements.

Fill in the blank: The electronic configuration of copper is [Ar] $3d^{10}$ _____.

2.2.3 Applications of electronic configurations in chemical behaviour

Electronic configurations explain many properties of elements, such as their valency, reactivity, and placement in the periodic table. For instance, alkali metals have one electron in their outermost s orbital, making them highly reactive. Noble gases have full outer shells, which makes them stable and chemically inert.

Understanding configurations also helps predict bonding patterns. For example, oxygen with six valence electrons tends to form two covalent bonds to achieve a stable octet.

Fill in the blank: Elements with full outer shells, such as neon, are chemically _____.

Pilot questions 2.2

1. Write the electronic configuration of carbon.
2. What is the electronic configuration of sodium?
3. Why does chromium have the configuration [Ar] $3d^5\ 4s^1$ instead of [Ar] $3d^4\ 4s^2$?
4. Which group of elements has full outer shells and is chemically inert?

5. How does the electronic configuration of oxygen explain its bonding behaviour?

2.3 Periodicity and Trends in the Periodic Table

2.3.1 Atomic radii and ionic radii

The **atomic radius** is the distance from the nucleus of an atom to the outermost electron. It provides a measure of the size of an atom. The **ionic radius** refers to the size of an ion, which changes depending on whether the atom loses or gains electrons. Cations (positively charged ions) are smaller than their parent atoms because they lose electrons, while anions (negatively charged ions) are larger because they gain electrons.

Across a period in the periodic table, atomic radii decrease due to increasing nuclear charge pulling electrons closer to the nucleus. Down a group, atomic radii increase because additional electron shells are added.

Fill in the blank: When an atom loses electrons to form a cation, its ionic radius becomes _____ than the atomic radius.

2.3.2 Ionisation energy and electron affinity

Ionisation energy is the amount of energy required to remove an electron from an atom in its gaseous state. It increases across a period due to stronger nuclear attraction and decreases down a group as electrons are farther from the nucleus.

Electron affinity is the energy change when an atom gains an electron. Elements with high electron affinity, such as halogens, readily accept electrons to form negative ions.

Fill in the blank: The energy required to remove an electron from an atom is called _____ energy.

2.3.3 Electronegativity and chemical reactivity

Electronegativity is the ability of an atom to attract electrons in a chemical bond. It increases across a period and decreases down a group. Fluorine is the most electronegative element.

Periodic trends in electronegativity explain chemical reactivity. For example, alkali metals are highly reactive because of their low electronegativity and tendency to lose electrons, while halogens are reactive due to their high electronegativity and tendency to gain electrons.

Fill in the blank: The most electronegative element in the periodic table is _____.

Pilot questions 2.3

1. Define atomic radius and explain how it changes across a period.
2. Why are cations smaller than their parent atoms?

3. What is ionisation energy?
4. Which group of elements has the highest electron affinity?
5. Name the most electronegative element in the periodic table.

2.4 Molecular Shapes and Hybridisation

2.4.1 Valence Shell Electron Pair Repulsion (VSEPR) theory

The **Valence Shell Electron Pair Repulsion (VSEPR) theory** explains the shapes of molecules based on the repulsion between electron pairs around a central atom. Electron pairs, whether bonding or non-bonding, arrange themselves as far apart as possible to minimise repulsion. This arrangement determines the geometry of the molecule.

For example, methane (CH_4) has four bonding pairs around carbon, which arrange themselves in a tetrahedral geometry. Water (H_2O), with two bonding pairs and two lone pairs, adopts a bent shape.

Fill in the blank: According to VSEPR theory, electron pairs around a central atom arrange themselves to minimise _____.

2.4.2 Common molecular geometries

Molecules adopt different geometries depending on the number of bonding and non-bonding pairs around the central atom. Some common geometries include:

- **Linear:** Two bonding pairs, 180° apart (e.g., CO_2).
- **Trigonal planar:** Three bonding pairs, 120° apart (e.g., BF_3).
- **Tetrahedral:** Four bonding pairs, 109.5° apart (e.g., CH_4).
- **Trigonal bipyramidal:** Five bonding pairs, with 90° and 120° angles (e.g., PCl_5).
- **Octahedral:** Six bonding pairs, 90° apart (e.g., SF_6).

Fill in the blank: A molecule with four bonding pairs around the central atom adopts a _____ geometry.

2.4.3 Hybridisation of orbitals and examples

Hybridisation is the process by which atomic orbitals mix to form new hybrid orbitals that are suitable for bonding. This concept explains molecular shapes and bond strengths.

- **sp hybridisation:** One s orbital mixes with one p orbital, forming two linear orbitals (e.g., $BeCl_2$).
- **sp^2 hybridisation:** One s orbital mixes with two p orbitals, forming three trigonal planar orbitals (e.g., BF_3).

- **sp^3 hybridisation:** One s orbital mixes with three p orbitals, forming four tetrahedral orbitals (e.g., CH_4).
- **sp^3d hybridisation:** One s, three p, and one d orbital mix to form trigonal bipyramidal orbitals (e.g., PCl_5).
- **sp^3d^2 hybridisation:** One s, three p, and two d orbitals mix to form octahedral orbitals (e.g., SF_6).

Fill in the blank: The hybridisation that explains the tetrahedral geometry of methane is _____.

Pilot questions 2.4

1. State the main idea of VSEPR theory.
2. What is the molecular geometry of carbon dioxide?
3. Give one example of a molecule with trigonal planar geometry.
4. Which geometry has bond angles of 109.5°?
5. What type of hybridisation explains the tetrahedral shape of methane?

Series Summary

This Series has focused on how the arrangement of electrons within atoms governs their properties, explains periodic trends, and determines the shapes of molecules. Its purpose was to build on the foundations of atomic theory and chemical reactions by showing how electronic structure connects directly to chemical behaviour and molecular geometry, making it a vital step in your study of chemistry.

We began with the modern electronic theory of atoms, exploring energy levels, quantum numbers, and the principles that guide electron arrangements. Then we examined how to write electronic configurations for both main group and transition elements, including notable exceptions, and considered their applications in predicting chemical behaviour. Following that, we studied periodicity, looking at trends such as atomic radii, ionisation energy, electron affinity, and electronegativity. Finally, we concluded with molecular shapes and hybridisation, using VSEPR theory to predict geometries and orbital mixing to explain bonding. In line with the Series Learning Outcomes, you should now be able to explain atomic structure, demonstrate electron configurations, analyse periodic trends, and apply theories of molecular geometry and hybridisation to chemical systems.

Concept Check Questions [Series 2]

Objective questions

1. The quantum number that defines the shape of an orbital is the _____ quantum number. (Linked to SLO 2.1)
2. Which principle states that electrons fill orbitals starting from the lowest energy level? (Linked to SLO 2.2)
3. The electronic configuration of sodium is _____. (Linked to SLO 2.3)
4. Across a period, atomic radius generally _____ due to increasing nuclear charge. (Linked to SLO 2.5)
5. The most electronegative element in the periodic table is _____. (Linked to SLO 2.5)

Short-answer questions

6. Define an orbital and give one example. (Linked to SLO 2.1)
7. State Hund's rule and explain its significance. (Linked to SLO 2.2)
8. Write the electronic configuration of oxygen. (Linked to SLO 2.3)
9. Explain why cations are smaller than their parent atoms. (Linked to SLO 2.5)
10. What is the molecular geometry of methane according to VSEPR theory? (Linked to SLO 2.6)

Long-answer questions

11. Analyse how electronic configurations influence the chemical behaviour of elements, using examples from alkali metals and noble gases. (Linked to SLO 2.4)
12. Demonstrate how VSEPR theory predicts molecular shapes, and evaluate its usefulness by comparing methane and water. (Linked to SLO 2.6)
13. Describe the concept of hybridisation and illustrate how it explains the bonding in methane and ethene. (Linked to SLO 2.7)

Answers to Concept Check Questions [Series 2]

Objective questions

1. Azimuthal quantum number.
2. Aufbau principle.
3. $1s^2\ 2s^2\ 2p^6\ 3s^1$.
4. Decreases.
5. Fluorine.

Short-answer questions

6. An orbital is a region in space where there is a high probability of finding an electron. Example: 2p orbital.

7. Hund's rule states that electrons occupy orbitals of equal energy singly with parallel spins before pairing. This minimises repulsion and stabilises the atom.

8. $1s^2\ 2s^2\ 2p^4$.

9. Cations are smaller because they lose electrons, reducing electron repulsion and allowing the nucleus to pull remaining electrons closer.

10. Tetrahedral geometry.

Long-answer questions

11. *Basic response*: Electronic configurations influence chemical behaviour by determining valency and reactivity. Alkali metals with one outer electron are highly reactive, while noble gases with full outer shells are inert. *Value-added response*: Beyond basic reactivity, electronic configurations explain trends in ionisation energy and bonding. For example, alkali metals readily form ionic compounds due to low ionisation energy, while noble gases are used in applications like lighting because of their stability.

12. *Basic response*: VSEPR theory predicts shapes by minimising electron pair repulsion. Methane has four bonding pairs around carbon, giving a tetrahedral shape, while water has two bonding pairs and two lone pairs, resulting in a bent shape. *Value-added response*: VSEPR theory is useful for predicting simple geometries but has limitations in complex molecules. For instance, it does not fully account for orbital hybridisation, which provides a deeper explanation of bond angles and strengths.

13. *Basic response*: Hybridisation is the mixing of atomic orbitals to form new orbitals for bonding. Methane involves sp^3 hybridisation, giving a tetrahedral shape, while ethene involves sp^2 hybridisation, giving a trigonal planar shape. *Value-added response*: Hybridisation explains not only shapes but also bond strengths and overlaps. In ethene, sp^2 hybridisation allows for strong sigma bonds and a pi bond, accounting for its reactivity in addition reactions.

Answers to Pilot Questions [Series 2]

Part 1: Modern Electronic Theory of Atoms

- Protons and neutrons.
- Principal, azimuthal, magnetic, and spin quantum numbers.
- Dumbbell-shaped.
- Electrons fill orbitals starting from the lowest energy level.
- Singly.

Part 2: Electronic Configurations of Elements

- $1s^2\ 2s^2\ 2p^2$.
- $1s^2\ 2s^2\ 2p^6\ 3s^1$.
- Because half-filled d orbitals provide extra stability.

- Noble gases.
- Oxygen tends to form two covalent bonds to complete its octet.

Part 3: Periodicity and Trends in the Periodic Table

- Atomic radius is the distance from the nucleus to the outermost electron; it decreases across a period.
- Because they lose electrons, reducing electron repulsion.
- The energy required to remove an electron from an atom in its gaseous state.
- Halogens.
- Fluorine.

Part 4: Molecular Shapes and Hybridisation

- Electron pairs arrange themselves to minimise repulsion.
- Linear geometry.
- BF_3.
- Tetrahedral geometry.
- sp^3 hybridisation.

Course code: CHM 101

Course Title: General Chemistry I

Course Unit: 2 Units

Series 3: Chemical Bonding, States of Matter, and Thermochemistry

Part 1 Chemical bonding

- Types of chemical bonds (ionic, covalent, metallic)
- Bond polarity and electronegativity
- Bond energy and bond length
- Intermolecular forces (hydrogen bonding, van der Waals forces)

Part 2 States of matter

- Properties of solids, liquids, and gases
- Gas laws (Boyle's, Charles's, Avogadro's, Ideal Gas Law)
- Real gases and deviations from ideal behaviour
- Phase changes and phase diagrams

Part 3 Thermochemistry

- Heat, work, and internal energy
- Enthalpy and calorimetry
- Hess's law and standard enthalpies of formation
- Bond enthalpies and energy changes in reactions

Introduction

In the last Series, we examined how electrons are arranged in atoms, how periodic trends arise, and how molecular shapes can be predicted. That knowledge provides the foundation for understanding how atoms interact with one another to form compounds, how matter exists in different physical states, and how energy changes accompany chemical processes. This Series will therefore connect microscopic bonding interactions with macroscopic properties and energetic considerations, making it central to your grasp of general chemistry.

The aim of this Series is to introduce you to the principles of chemical bonding, the behaviour of matter in its various states, and the fundamentals of thermochemistry. By the end, you should be able to explain how bonds form, describe the properties of solids, liquids, and gases, and analyse energy changes in chemical reactions.

We shall commence this Series by exploring chemical bonding, focusing on ionic, covalent, and metallic bonds, as well as bond polarity, bond energy, and intermolecular forces. Next, we will move on to the states of matter, considering the properties of solids, liquids, and gases, the gas laws, real gas behaviour, and phase changes. Finally, we will wrap up with thermochemistry, where we will study heat, work, enthalpy, calorimetry, Hess's law, and bond enthalpies, bringing the Series to a close with a clear understanding of how energy is involved in chemical reactions.

Series Learning Outcomes

Upon completing this Series, you should be able to:

- **SLO 3.1**: Define the different types of chemical bonds, including ionic, covalent, and metallic bonds.
- **SLO 3.2**: Explain the concepts of bond polarity, bond energy, bond length, and intermolecular forces.
- **SLO 3.3**: Describe the properties of solids, liquids, and gases, and distinguish between them.
- **SLO 3.4**: Apply the gas laws to calculate relationships among pressure, volume, temperature, and amount of gas.

- **SLO 3.5**: Analyse deviations of real gases from ideal behaviour using appropriate models.
- **SLO 3.6**: Demonstrate how to interpret phase changes and phase diagrams.
- **SLO 3.7**: Define heat, work, and internal energy in thermochemical processes.
- **SLO 3.8**: Calculate enthalpy changes using calorimetry and Hess's law.
- **SLO 3.9**: Evaluate energy changes in chemical reactions using bond enthalpies.

3.1 Chemical Bonding

3.1.1 Types of chemical bonds

A **chemical bond** is the force of attraction that holds atoms or ions together in a compound. The three main types are:

- **Ionic bonds:** These occur when electrons are transferred from one atom to another, usually between metals and non-metals. The resulting cations and anions are held together by electrostatic attraction.
- **Covalent bonds:** These involve the sharing of electron pairs between atoms, typically between non-metals. The shared electrons allow each atom to achieve a stable configuration.
- **Metallic bonds:** Found in metals, these bonds involve delocalised valence electrons moving freely through a lattice of positive ions, giving rise to properties such as conductivity and malleability.

Fill in the blank: A bond formed by the sharing of electron pairs between atoms is called a _____ bond.

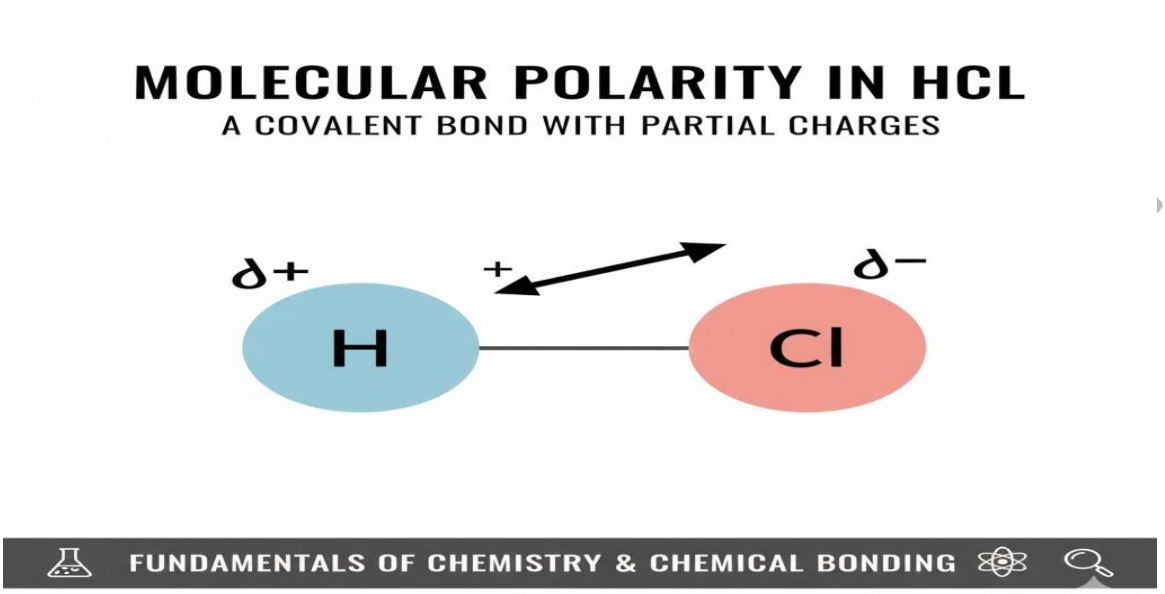
3.1.2 Bond polarity and electronegativity

Electronegativity is the ability of an atom to attract electrons in a chemical bond. When two atoms with different electronegativities form a covalent bond, the electrons are shared unequally, creating a **polar bond**. If the difference in electronegativity is very large, the bond may become ionic.

For example, in hydrogen chloride (HCl), chlorine is more electronegative than hydrogen, so the bond is polar with partial charges on each atom.

Fill in the blank: When electrons are shared unequally between atoms, the bond is described as _____

Plate 3.1: Polarity in a covalent bond



3.1.3 Bond energy and bond length

Bond energy is the amount of energy required to break one mole of a bond in the gaseous state. It reflects the strength of the bond. **Bond length** is the average distance between the nuclei of two bonded atoms. Stronger bonds generally have shorter bond lengths.

For example, a carbon–carbon triple bond (C≡C) has higher bond energy and shorter bond length than a carbon–carbon double bond (C=C).

Fill in the blank: The average distance between the nuclei of two bonded atoms is called bond _____.

Bond Type	Bond Energy (kJ/mol)	Bond Length (nm)
C–C	331	0.154
C=C	590	0.134
C≡C	812	0.120

Table 3.1: Bond energies and bond lengths of selected bonds

3.1.4 Intermolecular forces

Intermolecular forces are weak forces of attraction between molecules. They are not chemical bonds but strongly influence physical properties such as boiling and melting points.

- **Van der Waals forces:** Weak attractions due to temporary dipoles.

- **Dipole–dipole interactions:** Attractions between polar molecules.
- **Hydrogen bonding:** A strong type of dipole–dipole interaction involving hydrogen bonded to highly electronegative atoms such as oxygen, nitrogen, or fluorine.

For example, hydrogen bonding in water explains its unusually high boiling point compared to other molecules of similar size.

Fill in the blank: The unusually high boiling point of water is due to _____ bonding.

Pilot questions 3.1

1. What type of bond is formed when electrons are transferred between atoms?
2. Define electronegativity.
3. What is the difference between bond energy and bond length?
4. Name the three main types of intermolecular forces.
5. Why does water have a high boiling point compared to similar molecules?

3.2 States of Matter

3.2.1 Properties of solids, liquids, and gases

Matter exists in three main physical states: **solids, liquids, and gases**.

- **Solids:** Particles are closely packed in fixed positions, giving solids a definite shape and volume. They are rigid and incompressible.
- **Liquids:** Particles are close together but can move past one another, so liquids have a definite volume but take the shape of their container. They are less rigid and slightly compressible.
- **Gases:** Particles are far apart and move freely, so gases have neither definite shape nor volume. They are highly compressible and expand to fill their container.

Fill in the blank: A state of matter with definite volume but no fixed shape is called a _____.

3.2.2 Gas laws

The behaviour of gases can be described using several **gas laws**:

- **Boyle's law:** At constant temperature, the pressure of a gas is inversely proportional to its volume.
- **Charles's law:** At constant pressure, the volume of a gas is directly proportional to its temperature.
- **Avogadro's law:** Equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.
- **Ideal Gas Law:** Combines these relationships into the equation PV=nRT, where Pis pressure, Vis volume, nis moles of gas, Ris the gas constant, and Tis temperature.

Fill in the blank: The equation PV=nRTis known as the _____ Gas Law.

3.2.3 Real gases and deviations from ideal behaviour

The **Ideal Gas Law** assumes that gas particles have no volume and no intermolecular forces. However, real gases deviate from this behaviour, especially at high pressures and low temperatures.

The **van der Waals equation** modifies the Ideal Gas Law to account for particle volume and intermolecular forces:

$$(P+an2V2)(V-nb)=nRT$$

where aand bare constants specific to each gas.

Fill in the blank: Real gases deviate most from ideal behaviour at _____ pressure and _____ temperature.

3.2.4 Phase changes and phase diagrams

Phase changes are transitions between states of matter, such as melting, freezing, vaporisation, condensation, sublimation, and deposition. These processes involve energy changes but not chemical changes.

A **phase diagram** shows the conditions of temperature and pressure under which different states of matter exist. It includes the **triple point**, where solid, liquid, and gas coexist, and the **critical point**, beyond which a gas cannot be liquefied.

Fill in the blank: The point on a phase diagram where solid, liquid, and gas coexist is called the _____ point.

Pilot questions 3.2

1. What state of matter has definite shape and volume?
2. State Boyle's law in words.
3. Write the equation of the Ideal Gas Law.
4. Under what conditions do real gases deviate most from ideal behaviour?
5. What is the triple point in a phase diagram?

3.3 Thermochemistry

3.3.1 Heat, work, and internal energy

Thermochemistry is the study of energy changes, particularly heat, during chemical reactions. **Heat** is the transfer of energy due to a temperature difference, while **work** is energy transfer that results from a force acting through a distance. The **internal energy** of a system is the sum of all kinetic and potential energies of its particles.

In chemical reactions, energy can be absorbed or released. An endothermic process absorbs heat, while an exothermic process releases heat.

Fill in the blank: A chemical reaction that absorbs heat from its surroundings is called an _____ process.

3.3.2 Enthalpy and calorimetry

Enthalpy (H) is the heat content of a system at constant pressure. The change in enthalpy (H) indicates whether a reaction is endothermic ($H>0$) or exothermic ($H<0$).

Calorimetry is the experimental measurement of heat changes. A calorimeter is used to measure the heat absorbed or released during a chemical or physical process. For example, burning a fuel in a calorimeter allows us to determine its enthalpy of combustion.

Fill in the blank: The instrument used to measure heat changes in chemical reactions is called a _____.

3.3.3 Hess's law and standard enthalpies of formation

Hess's law states that the enthalpy change of a reaction is the same whether it occurs in one step or several steps. This principle allows us to calculate enthalpy changes indirectly.

Standard enthalpy of formation (Hf**)** is the enthalpy change when one mole of a compound is formed from its elements in their standard states. These values are tabulated and widely used in calculations.

Fill in the blank: The enthalpy change of a reaction is independent of the path taken, according to _____ law.

3.3.4 Bond enthalpies and energy changes in reactions

Bond enthalpy is the energy required to break one mole of a particular bond in the gaseous state. It provides a way to estimate the enthalpy change of a reaction by comparing the energy required to break bonds in reactants with the energy released when new bonds form in products.

For example, in the combustion of methane, C–H and O=O bonds are broken, while C=O and O–H bonds are formed. The difference in bond enthalpies gives the overall enthalpy change.

Fill in the blank: The energy required to break one mole of a bond in the gaseous state is called bond _____.

Pilot questions 3.3

1. Define internal energy.
2. What is the difference between endothermic and exothermic processes?
3. What instrument is used to measure heat changes in reactions?
4. State Hess's law.
5. What is bond enthalpy?

Series Summary

This series has explored how atoms bond together, how matter behaves in different physical states, and how energy changes accompany chemical reactions. Its purpose was to connect microscopic interactions with macroscopic properties and energetic principles, helping you see the direct link between bonding, states of matter, and thermochemistry in everyday chemical systems.

We began with chemical bonding, examining ionic, covalent, and metallic bonds, as well as bond polarity, bond energy, and intermolecular forces. We then moved on to the states of matter, considering the properties of solids, liquids, and gases, the gas laws, real gas behaviour, and phase diagrams. Finally, we concluded with thermochemistry, where we studied heat, work,

enthalpy, calorimetry, Hess's law, and bond enthalpies. In line with the Series Learning Outcomes, you should now be able to define and explain different types of bonds, apply gas laws, analyse deviations in real gases, interpret phase changes, calculate enthalpy changes, and evaluate energy transformations in chemical reactions.

Concept Check Questions [Series 3]

Objective questions

1. A bond formed by the transfer of electrons between atoms is called an _____ bond. (Linked to SLO 3.1)
2. The ability of an atom to attract electrons in a chemical bond is termed _____. (Linked to SLO 3.2)
3. At constant temperature, the pressure of a gas is inversely proportional to its volume. This statement describes _____ law. (Linked to SLO 3.4)
4. The point on a phase diagram where solid, liquid, and gas coexist is called the _____ point. (Linked to SLO 3.6)
5. The enthalpy change of a reaction is independent of the path taken, according to _____ law. (Linked to SLO 3.8)

Short-answer questions

6. Define metallic bonding and explain one property it accounts for in metals. (Linked to SLO 3.1)
7. Distinguish between bond energy and bond length. (Linked to SLO 3.2)
8. State Charles's law in words. (Linked to SLO 3.4)
9. Explain why real gases deviate from ideal behaviour at high pressure. (Linked to SLO 3.5)
10. What instrument is used to measure heat changes in chemical reactions? (Linked to SLO 3.7)

Long-answer questions

11. Analyse the role of intermolecular forces in determining physical properties, using water and carbon dioxide as examples. (Linked to SLO 3.2)
12. Demonstrate how to apply the Ideal Gas Law to calculate the volume of a gas at given conditions, and evaluate its limitations. (Linked to SLO 3.4, 3.5)
13. Describe Hess's law and illustrate its application with a worked example. (Linked to SLO 3.8)
14. Evaluate how bond enthalpies can be used to estimate reaction enthalpies, with reference to combustion reactions. (Linked to SLO 3.9)

Answers to Concept Check Questions [Series 3]

Objective questions

1. Ionic bond.
2. Electronegativity.
3. Boyle's law.
4. Triple point.
5. Hess's law.

Short-answer questions

6. Metallic bonding involves delocalised electrons moving freely among positive ions. It explains conductivity in metals.
7. Bond energy is the energy required to break one mole of a bond, while bond length is the average distance between nuclei of bonded atoms.
8. Charles's law states that at constant pressure, the volume of a gas is directly proportional to its temperature.
9. At high pressure, gas particles occupy significant volume and intermolecular forces become important, causing deviations.
10. Calorimeter.

Long-answer questions

11. *Basic response*: Intermolecular forces affect boiling and melting points. Water has hydrogen bonding, giving it a high boiling point, while carbon dioxide has weak van der Waals forces, making it a gas at room temperature. *Value-added response*: Intermolecular forces also influence solubility and biological processes. For example, hydrogen bonding stabilises DNA base pairs, while weak forces determine the behaviour of gases in the atmosphere.
12. *Basic response*: The Ideal Gas Law PV=nRTcan be used to calculate gas volume when pressure, temperature, and moles are known. *Value-added response*: While useful, the law assumes no intermolecular forces and negligible particle volume. At high pressures or low temperatures, real gases deviate, requiring corrections such as the van der Waals equation.
13. *Basic response*: Hess's law states that the enthalpy change of a reaction is the same regardless of the path taken. *Value-added response*: For example, the enthalpy of formation of CO_2 can be calculated indirectly by combining enthalpy changes of combustion reactions. This principle is widely applied in thermochemical cycles and industrial energy calculations.
14. *Basic response*: Bond enthalpies allow estimation of reaction enthalpies by comparing energy required to break bonds with energy released when new bonds form. *Value-added response*: In combustion reactions, such as burning methane, bond enthalpy calculations provide approximate values for energy released. These estimates are crucial in fuel efficiency studies and environmental impact assessments.

Answers to Pilot Questions [Series 3]

Part 3.1: Chemical Bonding

- Ionic bond.

- Electronegativity is the ability of an atom to attract electrons in a chemical bond.
- Bond energy is the energy required to break one mole of a bond, while bond length is the average distance between the nuclei of bonded atoms.
- Van der Waals forces, dipole–dipole interactions, and hydrogen bonding.
- Hydrogen bonding.

Part 3.2: States of Matter

- Solid.
- Boyle's law states that at constant temperature, the pressure of a gas is inversely proportional to its volume.
- PV=nRT.
- High pressure and low temperature.
- Triple point.

Part 3.3: Thermochemistry

- Internal energy is the sum of all kinetic and potential energies of the particles in a system.
- Endothermic processes absorb heat, while exothermic processes release heat.
- Calorimeter.
- Hess's law states that the enthalpy change of a reaction is the same regardless of the path taken.
- Bond enthalpy is the energy required to break one mole of a bond in the gaseous state.

Course code: CHM 101

Course title: General Chemistry I

Course credit load: 2 Units

Series 4: Kinetics, Thermodynamics, and Radioactivity in Chemistry

Part 1 Chemical equilibrium

- Dynamic nature of equilibrium
- Law of mass action and equilibrium constant (Kc, Kp)
- Le Chatelier's principle and factors affecting equilibrium
- Applications of equilibrium in industry (e.g., Haber process)

Part 2 Acids, bases, and salts

- Definitions of acids and bases (Arrhenius, Brønsted–Lowry, Lewis)

- Strength of acids and bases, pH and pOH calculations
- Neutralisation reactions and formation of salts
- Buffer solutions and their importance

Part 3 Electrochemistry

- Redox reactions and electrode potentials
- Electrochemical cells (galvanic and electrolytic)
- Standard electrode potentials and applications
- Electrolysis and industrial applications (e.g., electroplating, extraction of metals)

Introduction

In the last Series, we explored the foundational principles of chemical bonding, states of matter, and energy changes in reactions. Building on that knowledge, we now turn our attention to the dynamic processes that govern how reactions occur, how energy is transferred, and how unstable nuclei transform. These topics are central to understanding both everyday chemical phenomena and more advanced applications in fields such as medicine, engineering, and environmental science.

The aim of this Series is to deepen your understanding of chemical kinetics, thermodynamics, and radioactivity, and to equip you with the analytical tools needed to interpret reaction behaviour, energy flow, and nuclear transformations.

We shall commence this Series by examining the factors that influence reaction rates and how they are measured. Next, we will explore the principles of thermodynamics, focusing on energy changes, enthalpy, and entropy. Following that, we will investigate the nature of radioactive decay and its implications. To conclude, we will integrate these concepts to analyse how chemical systems evolve over time and how energy considerations shape their behaviour.

Series Learning Outcomes

Upon completing this Series, you should be able to:

- **SLO 4.1** define the concept of reaction rate and identify factors that influence it,
- **SLO 4.2** explain the role of activation energy and catalysts in chemical kinetics,
- **SLO 4.3** apply rate laws to calculate reaction order and rate constants,

- **SLO 4.4** describe the laws of thermodynamics and their implications for chemical systems,
- **SLO 4.5** calculate enthalpy, entropy, and Gibbs free energy changes for chemical reactions,
- **SLO 4.6** analyse the spontaneity of reactions using thermodynamic parameters,
- **SLO 4.7** define radioactivity and distinguish between alpha, beta, and gamma radiation,
- **SLO 4.8** demonstrate how to balance nuclear equations for radioactive decay processes, and
- **SLO 4.9** evaluate practical applications and hazards of radioactivity in medicine, industry, and the environment.

4.1 Chemical Equilibrium

4.1.1 Dynamic nature of equilibrium

Chemical equilibrium is the state in which the rate of the forward reaction equals the rate of the reverse reaction, so the concentrations of reactants and products remain constant over time. It is important to note that equilibrium is **dynamic**, meaning reactions continue to occur in both directions, but with no net change in composition.

For example, in the synthesis of ammonia from nitrogen and hydrogen, molecules of ammonia are constantly being formed and decomposed, yet the overall concentrations remain steady once equilibrium is reached.

Fill in the blank: At equilibrium, the rate of the forward reaction is _____ to the rate of the reverse reaction.

4.1.2 Law of mass action and equilibrium constant

The **law of mass action** states that the rate of a chemical reaction is proportional to the product of the concentrations of the reactants, each raised to a power equal to its stoichiometric coefficient.

The **equilibrium constant (Kc)** is the ratio of the concentrations of products to reactants at equilibrium, each raised to their stoichiometric coefficients. For gaseous reactions, the equilibrium constant can also be expressed in terms of partial pressures as **Kp**.

For example, for the reaction:

$$N2(g)+3H2(g)\rightleftharpoons 2NH3(g)$$

$$Kc=[NH3]2[N2][H2]3$$

Fill in the blank: The equilibrium constant expressed in terms of partial pressures is denoted as _____.

Reaction Type	Example Reaction	Equilibrium Constant Expression
General reaction	(aA + bB \rightleftharpoons cC + dD)	(K = \dfrac{[C]^c [D]^d}{[A]^a [B]^b})
Synthesis (formation)	(N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g))	(K = \dfrac{[NH_3]^2}{[N_2][H_2]^3})
Decomposition	(2HI(g) \rightleftharpoons H_2(g) + I_2(g))	(K = \dfrac{[H_2][I_2]}{[HI]^2})
Neutralisation	(HCl(aq) + NaOH(aq) \rightleftharpoons NaCl(aq) + H_2O(l))	(K = \dfrac{[NaCl]}{[HCl][NaOH]}) (water omitted as pure liquid)
Precipitation	(AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq))	(K_{sp} = [Ag^+][Cl^-])
Gas-phase equilibrium (using partial pressures)	(2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g))	(K_p = \dfrac{(P_{SO_3})^2}{(P_{SO_2})^2 (P_{O_2})})

Table 4.1: Examples of equilibrium constant expressions

4.1.3 Le Chatelier's principle and factors affecting equilibrium

Le Chatelier's principle states that if a system at equilibrium is disturbed by a change in concentration, temperature, or pressure, the system will shift to counteract the disturbance and restore equilibrium.

4.1.4 Applications of equilibrium in industry

Chemical equilibrium principles are widely applied in industry to maximise yields. A classic example is the **Haber process**, where nitrogen and hydrogen are combined under high pressure and moderate temperature with a catalyst to produce ammonia.

By manipulating conditions such as pressure, temperature, and concentration, industries optimise production while balancing cost and efficiency.

Fill in the blank: The industrial process for synthesising ammonia from nitrogen and hydrogen is called the _____ process.

Pilot questions 4.1

1. What happens to the rates of forward and reverse reactions at equilibrium?
2. Write the equilibrium constant expression for the reaction $H_2 + I_2 \rightleftharpoons 2HI$.
3. State Le Chatelier's principle.

4. How does increasing temperature affect an endothermic equilibrium reaction?
5. Name one industrial process that applies equilibrium principles.

4.2 Acids, Bases, and Salts

4.2.1 Definitions of acids and bases

An **acid** is a substance that donates protons (H^+ ions) or increases hydrogen ion concentration in solution. A **base** is a substance that accepts protons or increases hydroxide ion concentration in solution.

There are three main definitions:

- **Arrhenius definition:** Acids produce H^+ in water, bases produce OH^-.
- **Brønsted–Lowry definition:** Acids are proton donors, bases are proton acceptors.
- **Lewis definition:** Acids accept electron pairs, bases donate electron pairs.

Fill in the blank: According to Brønsted–Lowry, an acid is a proton _____.

4.2.2 Strength of acids and bases, pH and pOH calculations

Strong acids and **strong bases** ionise completely in solution, while weak acids and bases ionise only partially. The **pH** scale measures hydrogen ion concentration, defined as:

$$pH=-\log[H+]$$

Similarly,

$$pOH=-\log[OH-]$$

and

$$pH+pOH=14$$

Fill in the blank: A solution with pH less than 7 is _____.

4.2.3 Neutralisation reactions and formation of salts

A **neutralisation reaction** occurs when an acid reacts with a base to form a **salt** and water. For example:

$$HCl+NaOHNaCl+H2O$$

Salts can be neutral, acidic, or basic depending on the strength of the acid and base from which they are formed.

Fill in the blank: The product of a neutralisation reaction is a salt and _____.

4.2.4 Buffer solutions and their importance

A **buffer solution** is a solution that resists changes in pH when small amounts of acid or base are added. Buffers usually consist of a weak acid and its conjugate base, or a weak base and its conjugate acid.

For example, a mixture of acetic acid (CH_3COOH) and sodium acetate (CH_3COONa) acts as a buffer. Buffers are essential in biological systems, such as maintaining blood pH around 7.4.

Fill in the blank: A buffer solution resists changes in _____ when small amounts of acid or base are added.

Pilot questions 4.2

1. State the Arrhenius definition of an acid.
2. What is the pH of a solution with $[H^+] = 1 \times 10^{-3}$ M?
3. Write the general equation for a neutralisation reaction.
4. Give one example of a buffer solution.
5. Why are buffers important in biological systems?

4.3 Electrochemistry

4.3.1 Redox reactions and electrode potentials

Electrochemistry is the study of chemical processes that involve the movement of electrons. A **redox reaction** is a chemical reaction in which oxidation (loss of electrons) and reduction (gain of electrons) occur simultaneously.

In electrochemical systems, the tendency of a species to gain or lose electrons is measured as its **electrode potential**, expressed relative to the standard hydrogen electrode (SHE). Positive potentials indicate a greater tendency to be reduced, while negative potentials indicate a greater tendency to be oxidised.

Fill in the blank: The electrode used as a reference with zero potential is the _____ hydrogen electrode.

4.3.2 Electrochemical cells (galvanic and electrolytic)

An **electrochemical cell** is a device that converts chemical energy into electrical energy or vice versa.

- **Galvanic (voltaic) cells:** These generate electricity from spontaneous redox reactions. A common example is the Daniell cell, which uses zinc and copper electrodes.
- **Electrolytic cells:** These use electrical energy to drive non-spontaneous reactions, such as electrolysis of water.

Fill in the blank: A cell that generates electricity from a spontaneous redox reaction is called a _____ cell.

4.3.3 Standard electrode potentials and applications

Standard electrode potential (E°) is the potential of a half-cell measured under standard conditions (1 M concentration, 1 atm pressure, 25 °C). These values are tabulated and used to predict the direction of redox reactions.

The overall cell potential is calculated as:

$$E_{cell} = E_{cathode} - E_{anode}$$

Applications include predicting spontaneity of reactions, designing batteries, and understanding corrosion.

Fill in the blank: The overall cell potential is calculated as E°cell = E°cathode – E°_____.

4.3.4 Electrolysis and industrial applications

Electrolysis is the process of using electrical energy to drive a non-spontaneous chemical reaction. It is widely used in industry.

Examples include:

- **Electroplating:** Depositing a thin layer of metal onto an object.
- **Extraction of metals:** Obtaining aluminium from bauxite using electrolysis of molten aluminium oxide.
- **Production of chemicals:** Electrolysis of brine to produce chlorine, hydrogen, and sodium hydroxide.

  Fill in the blank: The industrial extraction of aluminium from bauxite involves _____ of molten aluminium oxide.

Pilot questions 4.3

1. What is a redox reaction?
2. Name the reference electrode with zero potential.
3. Distinguish between galvanic and electrolytic cells.
4. Write the formula for calculating overall cell potential.
5. Give one industrial application of electrolysis.

Series Summary

This Series has focused on three interconnected areas of chemistry: kinetics, thermodynamics, and radioactivity. Its purpose was to help you understand how chemical reactions proceed, the energy changes that accompany them, and the behaviour of unstable nuclei. These topics are highly relevant, as they explain both everyday chemical processes and advanced applications in industry, medicine, and environmental science.

We began with kinetics, examining reaction rates, activation energy, catalysts, and the use of rate laws. We then moved on to thermodynamics, where we explored the laws governing energy, calculated enthalpy, entropy, and Gibbs free energy, and analysed reaction spontaneity. Finally, we concluded with radioactivity, defining types of radiation, balancing nuclear equations, and evaluating practical uses and hazards. In line with the Series Learning Outcomes, you should now be able to define and explain reaction rates, apply thermodynamic principles to chemical systems, analyse spontaneity, and evaluate both the applications and risks of radioactivity.

Concept Check Questions [Series 4]

Objective questions

1. The minimum energy required for a chemical reaction to occur is called _____. (Linked to SLO 4.2)
2. The rate of a chemical reaction is proportional to the product of reactant concentrations. This statement describes the _____ law. (Linked to SLO 4.3)
3. The First Law of Thermodynamics states that energy cannot be _____ or destroyed. (Linked to SLO 4.4)
4. A reaction is spontaneous if Gibbs free energy change (ΔG) is _____. (Linked to SLO 4.6)
5. The emission of an alpha particle reduces the mass number of a nucleus by _____. (Linked to SLO 4.7)

Short-answer questions

6. Define reaction rate and name two factors that affect it. (Linked to SLO 4.1)
7. Explain the role of a catalyst in chemical kinetics. (Linked to SLO 4.2)
8. Calculate the pH of a solution where $[H^+] = 1 \times 10^{-4}$ M. (Linked to SLO 4.5)
9. Distinguish between alpha and beta radiation. (Linked to SLO 4.7)
10. Write a balanced nuclear equation for the beta decay of carbon-14. (Linked to SLO 4.8)

Long-answer questions

11. Analyse how temperature and concentration affect reaction rates, using collision theory to support your explanation. (Linked to SLO 4.1, 4.3)
12. Demonstrate how to calculate Gibbs free energy change for a reaction, and evaluate its significance in predicting spontaneity. (Linked to SLO 4.5, 4.6)
13. Describe the three laws of thermodynamics and illustrate their importance in chemical systems. (Linked to SLO 4.4, 4.6)

14. Evaluate the applications and hazards of radioactivity in medicine and industry, giving specific examples. (Linked to SLO 4.9)

Answers to Concept Check Questions [Series 4]

Objective questions

1. Activation energy.
2. Law of mass action.
3. Created.
4. Negative.
5. Four.

Short-answer questions

6. Reaction rate is the change in concentration of reactants or products per unit time. Factors include temperature and concentration.
7. A catalyst lowers activation energy, increasing the rate of reaction without being consumed.
8. $pH = -\log(1 \times 10^{-4}) = 4$.
9. Alpha radiation involves emission of a helium nucleus (2 protons, 2 neutrons), while beta radiation involves emission of an electron or positron.
10. 614C714N+-.

Long-answer questions

11. *Basic response*: Higher temperature increases particle kinetic energy, leading to more frequent and energetic collisions. Higher concentration increases collision frequency, both raising reaction rate. *Value-added response*: Collision theory explains that only collisions with sufficient energy and correct orientation lead to reaction. Catalysts provide alternative pathways with lower activation energy, further enhancing rates.
12. *Basic response*: Gibbs free energy is calculated as $\Delta G = \Delta H - T\Delta S$. A negative ΔG indicates spontaneity. *Value-added response*: For example, in the combustion of glucose, ΔG is highly negative, showing spontaneity. Gibbs free energy also connects chemical reactions to biological processes such as ATP hydrolysis.
13. *Basic response*: The First Law states energy is conserved. The Second Law states entropy increases in spontaneous processes. The Third Law states entropy of a perfect crystal at absolute zero is zero. *Value-added response*: These laws underpin energy transfer in engines, predict feasibility of reactions, and explain why absolute zero cannot be reached.
14. *Basic response*: Radioactivity is used in medical imaging and cancer treatment, but poses hazards such as radiation sickness. *Value-added response*: Applications include PET scans, radiotherapy, and industrial radiography. Hazards involve environmental contamination and long-term health risks, requiring strict safety protocols and regulation.

Answers to Pilot Questions [Series 4]

Part 4.1: Chemical Equilibrium

- Equal.
- HI]2/[H2][I2.
- If a system at equilibrium is disturbed, it shifts to oppose the change.
- It shifts towards products.
- Haber process.

Part 4.2: Acids, Bases, and Salts

- An acid produces H^+ ions in water.
- pH = 3.
- Acid + Base → Salt + Water.
- Acetic acid and sodium acetate.
- They maintain stable pH, essential for processes like blood regulation.

Part 4.3: Electrochemistry

- A reaction involving simultaneous oxidation and reduction.
- Standard hydrogen electrode.
- Galvanic cells generate electricity spontaneously; electrolytic cells require external energy.
- Ecell=Ecathode-Eanode.
- Electroplating or aluminium extraction**.**

Course code: CHM 101

Course Title: General Chemistry I

Course Unit: 2 Units

Series 5: Kinetics, Thermodynamics, and Radioactivity in Chemistry

Part 1 Chemical kinetics

- Definition of reaction rate

- Factors affecting reaction rate (concentration, temperature, catalysts, surface area)
- Collision theory and activation energy
- Rate laws and reaction order

Part 2 Thermodynamics

- First, Second, and Third Laws of Thermodynamics
- Enthalpy, entropy, and Gibbs free energy
- Spontaneity of reactions and equilibrium connections
- Applications of thermodynamic principles in chemical systems

Part 3 Radioactivity

- Nature of radioactivity and types of radiation (alpha, beta, gamma)
- Nuclear equations and half-life
- Applications of radioactivity (medicine, industry, energy)
- Hazards and safety measures

Introduction

Chemical reactions are not only about what substances form but also about how fast they occur, how much energy they involve, and what happens when nuclei themselves change. These aspects are captured in the study of kinetics, thermodynamics, and radioactivity. Together, they provide a deeper understanding of why reactions proceed as they do, how energy governs chemical systems, and how nuclear processes shape both natural phenomena and technological applications. This Series is therefore central to connecting the microscopic behaviour of particles with the macroscopic outcomes we observe in laboratories, industries, and everyday life.

The aim of this Series is to equip you with the ability to analyse reaction rates, apply thermodynamic principles to chemical systems, and evaluate the nature and applications of radioactivity. By the end, you should be able to calculate reaction parameters, assess spontaneity, and interpret nuclear transformations with confidence.

We shall commence this Series by exploring chemical kinetics, focusing on reaction rates, factors that influence them, and the role of activation energy and catalysts. Next, we will move on to thermodynamics, where we will examine the fundamental laws, enthalpy, entropy, and Gibbs free energy, and use these to predict spontaneity and equilibrium behaviour. Finally, we will wrap up with radioactivity, considering the types of radiation, nuclear equations, half-life, and both the applications and hazards of radioactive processes. This progression will round off the Series with a comprehensive view of how chemical and nuclear changes are governed by both speed and energy.

Series Learning Outcomes

Upon completing this Series, you should be able to:

- **SLO 5.1** define reaction rate and identify the main factors that influence it,
- **SLO 5.2** explain the concept of activation energy and the role of catalysts in chemical kinetics,
- **SLO 5.3** apply collision theory to analyse how particle interactions affect reaction rates,
- **SLO 5.4** calculate reaction order and rate constants using rate laws,
- **SLO 5.5** describe the First, Second, and Third Laws of Thermodynamics and their implications for chemical systems,
- **SLO 5.6** calculate enthalpy, entropy, and Gibbs free energy changes for chemical reactions,
- **SLO 5.7** analyse the spontaneity of reactions using thermodynamic parameters,
- **SLO 5.8** define radioactivity and distinguish between alpha, beta, and gamma radiation,
- **SLO 5.9** demonstrate how to balance nuclear equations for radioactive decay processes, and
- **SLO 5.10** evaluate practical applications and hazards of radioactivity in medicine, industry, and the environment.

5.1 Chemical Kinetics

5.1.1 Definition of reaction rate

Reaction rate is the change in concentration of reactants or products per unit time. It tells us how quickly a chemical reaction proceeds. Rates can be expressed in terms of the disappearance of reactants or the appearance of products.

For example, in the reaction:

$$2H2+O22H2O$$

the rate can be measured by how fast hydrogen decreases or how fast water increases.

Fill in the blank: Reaction rate measures the change in _____ of reactants or products per unit time.

5.1.2 Factors affecting reaction rate

Several factors influence how fast a reaction occurs:

- **Concentration**: Higher concentration increases collision frequency.

- **Temperature**: Higher temperature increases particle energy, leading to more effective collisions.
- **Catalysts**: A catalyst is a substance that speeds up a reaction by lowering activation energy without being consumed.
- **Surface area**: Greater surface area of reactants increases exposure and collision opportunities.

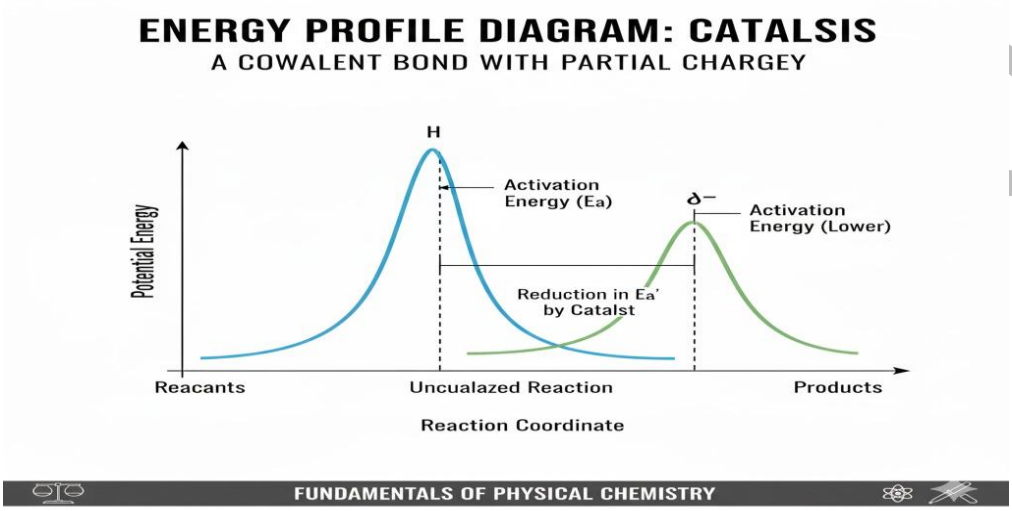


Plate 5.1: Energy profile with and without a catalyst

Fill in the blank: A catalyst increases reaction rate by lowering _____ energy.

5.1.3 Collision theory and activation energy

Collision theory states that chemical reactions occur when particles collide with sufficient energy and proper orientation. The minimum energy required for a reaction to occur is called **activation energy**.

Not all collisions lead to reactions. Only those with energy equal to or greater than the activation energy, and with the correct orientation, are successful.

Fill in the blank: According to collision theory, particles must collide with sufficient energy and correct _____ to react.

5.1.4 Rate laws and reaction order

A **rate law** expresses the relationship between reaction rate and the concentration of reactants. It is written as:

$$\text{Rate} = k[A]^m[B]^n$$

where k is the rate constant, and m and n are the reaction orders with respect to reactants A and B.

The **reaction order** indicates how the rate depends on the concentration of a reactant. It is determined experimentally, not from the balanced equation.

Fill in the blank: The constant in a rate law that depends on temperature and reaction conditions is called the _____ constant.

Pilot questions 5.1

1. Define reaction rate.
2. Name two factors that affect reaction rate.
3. State collision theory.
4. What is activation energy?
5. Write the general form of a rate law.

5.2 Thermodynamics

5.2.1 First, Second, and Third Laws of Thermodynamics

The **First Law of Thermodynamics** states that energy cannot be created or destroyed, only transformed from one form to another. This is essentially the law of conservation of energy applied to chemical systems.

The **Second Law of Thermodynamics** states that in any spontaneous process, the total entropy of the universe increases. **Entropy** is a measure of disorder or randomness in a system.

The **Third Law of Thermodynamics** states that the entropy of a perfect crystal at absolute zero is zero. This provides a reference point for calculating entropy values.

Fill in the blank: The entropy of a perfect crystal at absolute zero is _____.

5.2.2 Enthalpy, entropy, and Gibbs free energy

Enthalpy (ΔH) is the heat content of a system at constant pressure. **Entropy (ΔS)** measures the disorder or randomness of a system. **Gibbs free energy (ΔG)** combines enthalpy and entropy to predict spontaneity of reactions:

$$G=H-TS$$

A negative ΔG indicates a spontaneous reaction, while a positive ΔG indicates a non-spontaneous reaction.

Fill in the blank: A reaction is spontaneous if ΔG is _____.

5.2.3 Spontaneity of reactions and equilibrium connections

Spontaneity refers to whether a reaction occurs naturally without external input. A negative ΔG means the reaction is thermodynamically favourable.

At equilibrium, ΔG = 0, meaning there is no net change in the system. The relationship between Gibbs free energy and the equilibrium constant (K) is given by:

$$G=-RTln K$$

where R is the gas constant and T is temperature in Kelvin.

Fill in the blank: At equilibrium, the Gibbs free energy change (ΔG) is equal to _____.

5.2.4 Applications of thermodynamic principles in chemical systems

Thermodynamic principles are applied in many areas of chemistry and industry:

- Predicting whether reactions will occur spontaneously.
- Designing energy-efficient processes in chemical engineering.
- Understanding biological processes such as ATP hydrolysis.
- Explaining why certain reactions require external energy input.

Fill in the blank: The hydrolysis of ATP in biological systems is an example of a reaction predicted by _____ free energy.

Pilot questions 5.2

1. State the First Law of Thermodynamics.
2. What does entropy measure?
3. Write the Gibbs free energy equation.
4. What is the condition for spontaneity in terms of ΔG?
5. Give one application of thermodynamics in biological systems.

5.3 Radioactivity

5.3.1 Nature of radioactivity and types of radiation

Radioactivity is the spontaneous emission of particles or energy from unstable atomic nuclei. This occurs because certain nuclei have an imbalance of protons and neutrons, making them unstable.

There are three main types of radiation:

- **Alpha radiation (α):** Emission of a helium nucleus (2 protons and 2 neutrons).
- **Beta radiation (β):** Emission of an electron (β^-) or positron (β^+).

- **Gamma radiation (γ):** Emission of high-energy electromagnetic waves.

Fill in the blank: Alpha radiation involves the emission of a _____ nucleus.

5.3.2 Nuclear equations and half-life

A **nuclear equation** represents radioactive decay, showing the parent nucleus, emitted particle, and daughter nucleus. For example:

$$92238U90234Th+24He$$

Half-life is the time required for half of the radioactive nuclei in a sample to decay. It is a constant property of each isotope and is used to date materials and predict activity.

Fill in the blank: The time required for half of the radioactive nuclei in a sample to decay is called _____.

5.3.3 Applications of radioactivity

Radioactivity has many practical uses:

- **Medicine:** Radiotherapy for cancer treatment, diagnostic imaging (PET scans).
- **Industry:** Radiography for detecting structural flaws, sterilisation of equipment.
- **Energy:** Nuclear power generation using controlled fission reactions.

Fill in the blank: The medical use of radioactivity to treat cancer is called _____ therapy.

5.3.4 Hazards and safety measures

While radioactivity has many benefits, it also poses hazards. Exposure to high levels of radiation can damage living tissues, cause mutations, or lead to radiation sickness.

Safety measures include:

- Shielding with lead or concrete.
- Limiting exposure time.
- Maintaining safe distance from sources.
- Using protective clothing and monitoring devices.

Fill in the blank: One way to reduce radiation exposure is to use protective _____.

Pilot questions 5.3

1. Define radioactivity.
2. Name the three types of radiation.

3. What is half-life?
4. Give one medical application of radioactivity.
5. State one safety measure used when handling radioactive materials.

Series Summary

This Series has explored the essential principles of kinetics, thermodynamics, and radioactivity, all of which are central to understanding how chemical and nuclear processes occur. Its purpose was to help you connect the speed of reactions, the energy changes that govern them, and the behaviour of unstable nuclei to practical applications in science, industry, and everyday life.

We began with kinetics, where you examined reaction rates, the influence of factors such as temperature and catalysts, and the application of collision theory and rate laws. Then we moved on to thermodynamics, considering the three fundamental laws, enthalpy, entropy, and Gibbs free energy, and how these determine spontaneity and equilibrium. Finally, we concluded with radioactivity, defining types of radiation, balancing nuclear equations, and evaluating both applications and hazards. In line with the Series Learning Outcomes, you should now be able to define and analyse reaction rates, apply thermodynamic principles to predict spontaneity, and evaluate the uses and risks of radioactivity with confidence.

Concept Check Questions [Series 5]

Objective questions

1. The minimum energy required for a chemical reaction to occur is called _____. (Linked to SLO 5.2)
2. The theory that explains reaction rates based on particle collisions is called _____ theory. (Linked to SLO 5.3)
3. The First Law of Thermodynamics states that energy cannot be _____ or destroyed. (Linked to SLO 5.5)
4. A reaction is spontaneous if Gibbs free energy change (ΔG) is _____. (Linked to SLO 5.7)
5. The emission of a helium nucleus from an unstable atom is known as _____ radiation. (Linked to SLO 5.8)

Short-answer questions

6. Define reaction rate and name two factors that affect it. (Linked to SLO 5.1)
7. Explain the role of a catalyst in chemical kinetics. (Linked to SLO 5.2)

8. Write the Gibbs free energy equation. (Linked to SLO 5.6)
9. Distinguish between alpha and gamma radiation. (Linked to SLO 5.8)
10. Write a balanced nuclear equation for the alpha decay of uranium-238. (Linked to SLO 5.9)

Long-answer questions

11. Analyse how temperature and concentration affect reaction rates, using collision theory to support your explanation. (Linked to SLO 5.1, 5.3)
12. Demonstrate how to calculate Gibbs free energy change for a reaction, and evaluate its significance in predicting spontaneity. (Linked to SLO 5.6, 5.7)
13. Describe the three laws of thermodynamics and illustrate their importance in chemical systems. (Linked to SLO 5.5, 5.7)
14. Evaluate the applications and hazards of radioactivity in medicine, industry, and the environment, giving specific examples. (Linked to SLO 5.10)

Answers to Concept Check Questions [Series 5]

Objective questions

1. Activation energy.
2. Collision theory.
3. Created.
4. Negative.
5. Alpha radiation.

Short-answer questions

6. Reaction rate is the change in concentration of reactants or products per unit time. Factors include temperature and concentration.
7. A catalyst lowers activation energy, increasing the rate of reaction without being consumed.
8. $\Delta G = \Delta H - T\Delta S$.
9. Alpha radiation involves emission of a helium nucleus, while gamma radiation involves emission of high-energy electromagnetic waves.
10. $^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He$.

Long-answer questions

11. *Basic response*: Higher temperature increases particle kinetic energy, leading to more frequent and energetic collisions. Higher concentration increases collision frequency, both raising reaction rate. *Value-added response*: Collision theory explains that only collisions with sufficient energy and correct orientation lead to reaction. Catalysts provide alternative pathways with lower activation energy, further enhancing rates.
12. *Basic response*: Gibbs free energy is calculated as $\Delta G = \Delta H - T\Delta S$. A negative ΔG indicates spontaneity. *Value-added response*: For example, in the combustion of glucose, ΔG is

highly negative, showing spontaneity. Gibbs free energy also connects chemical reactions to biological processes such as ATP hydrolysis.

13. *Basic response*: The First Law states energy is conserved. The Second Law states entropy increases in spontaneous processes. The Third Law states entropy of a perfect crystal at absolute zero is zero. *Value-added response*: These laws underpin energy transfer in engines, predict feasibility of reactions, and explain why absolute zero cannot be reached.

14. *Basic response*: Radioactivity is used in medical imaging and cancer treatment, but poses hazards such as radiation sickness. *Value-added response*: Applications include PET scans, radiotherapy, and industrial radiography. Hazards involve environmental contamination and long-term health risks, requiring strict safety protocols and regulation.

Answers to Pilot Questions [Series 5]

Part 5.1: Chemical Kinetics

- Change in concentration.
- Concentration and temperature.
- Particles must collide with sufficient energy and correct orientation.
- Activation energy.
- Rate = $k[A]^{m}[B]^{n}$.

Part 5.2: Thermodynamics

- Energy cannot be created or destroyed, only transformed.
- Disorder or randomness.
- $\Delta G = \Delta H - T\Delta S$.
- Negative.
- ATP hydrolysis.

Part 5.3: Radioactivity

- Spontaneous emission of particles or energy from unstable nuclei.
- Alpha, beta, and gamma radiation.
- Half-life.
- Radiotherapy.
- Protective clothing.