Class 9 · Punjab Boards | Federal Board (PCTB) · English Medium

# Chemistry 9th Class

Complete Notes

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Prepared by freebooks.pk — original student notes covering the Punjab Curriculum and Textbook Board (PTB/PCTB) syllabus. Each chapter includes key concepts, definitions, formulas, diagrams, solved examples, short and long questions, MCQs, revision and exam tips.

# Unit 1: Fundamentals of Chemistry

Chemistry is the branch of science that deals with the composition, structure, properties and reactions of matter. This unit introduces the eight major branches of chemistry, then builds the basic vocabulary used throughout the subject: matter, substance, mixture, element, compound, atomic number, mass number, relative atomic mass, chemical formulae, ions, and molecules.

The second half of the unit introduces the mole concept -- the link between the mass of a substance and the number of particles (atoms, molecules or formula units) it contains -- along with Avogadro's Number, and the chemical calculations used to convert between mass, moles, and number of particles.

## Learning Objectives

- Identify and differentiate among the eight branches of chemistry
- Distinguish between matter, substance and mixture, and between physical and chemical properties
- Differentiate among elements, compounds and mixtures, and between homogeneous and heterogeneous mixtures
- Define atomic number, mass number, atomic mass unit and relative atomic mass based on the C-12 scale
- Differentiate between empirical formula and molecular formula, and between formula unit and molecular formula
- Define ions, cations, anions, molecular ions, formula units and free radicals, and distinguish among them
- Relate gram atomic mass, gram molecular mass and gram formula mass to the mole
- Describe how Avogadro's Number is related to a mole of any substance, and use it in chemical calculations

## Key Concepts

### 1.1 Branches of Chemistry

Chemistry is divided into eight main branches. Physical Chemistry deals with the relationship between the composition and physical properties of matter, such as the structure of atoms, the behaviour of gases, liquids and solids, and the effect of temperature or radiation on matter. Organic Chemistry studies covalent compounds of carbon and hydrogen (hydrocarbons) and their derivatives, covering petroleum, petrochemicals and pharmaceutical industries. Inorganic Chemistry studies all elements and their compounds except hydrocarbons and their derivatives, with applications in glass, cement, ceramics and metallurgy.

Biochemistry studies the structure, composition and chemical reactions of substances found in living organisms, including the synthesis and metabolism of carbohydrates, proteins and fats. Industrial Chemistry deals with manufacturing chemical compounds on a commercial scale, such as oxygen, chlorine, ammonia, caustic soda, nitric acid and sulphuric acid. Nuclear Chemistry deals with radioactivity, nuclear processes and atomic energy, with applications in radiotherapy, food preservation and nuclear power generation. Environmental Chemistry studies the components of the environment and the effects of human activity on it. Analytical Chemistry deals with the separation and analysis of a sample to identify its components -- qualitative analysis identifies what is present, while quantitative analysis determines how much of each component is present.

### 1.2 Basic Definitions: Matter, Substance and Mixture

Matter is anything that has mass and occupies space, and can exist in one of three physical states: solid, liquid or gas. A pure piece of matter with a fixed composition is called a substance, while impure matter with no fixed composition is called a mixture (homogeneous or heterogeneous).

Physical properties are associated with the physical state of a substance -- colour, smell, taste, hardness, solubility, melting point and boiling point -- and do not change its chemical composition (e.g. ice melting to water). Chemical properties depend on the composition of the substance; when a substance undergoes a chemical change, its composition changes and new substances form (e.g. water decomposing into hydrogen and oxygen gas).

### 1.2.1 Elements, Compounds and Mixtures

An element is a substance made up of the same type of atoms, having the same atomic number, and cannot be decomposed into simpler substances by ordinary chemical means. 118 elements are currently known, of which 92 occur naturally; about 80 percent of all elements are metals. Elements are represented by symbols -- a single capital letter (H, N, C) or a capital plus lowercase letter (Ca, Na, Cl) -- taken from the element's English, Latin, Greek or German name.

Valency is the combining capacity of an element, based on the number of electrons in its outermost shell. Elements with fewer than four valence electrons (like Na, Mg, Al) lose electrons to complete their octet, showing valencies of 1, 2 and 3 respectively; elements with five or more valence electrons (like N, O, Cl) gain electrons instead, showing valencies of 3, 2 and 1 respectively. Some elements, such as iron, show variable valency ($Fe^{2+}$ in ferrous compounds, $Fe^{3+}$ in ferric compounds).

A compound is a substance made of two or more elements chemically combined in a fixed ratio by mass; the elements lose their individual properties and the compound has entirely new properties (e.g. water forms from hydrogen and oxygen combining in a fixed 1:8 mass ratio). Ionic compounds form a three-dimensional crystal lattice and are represented by formula units (e.g. NaCl); covalent compounds mostly exist as molecules and are represented by molecular formulae (e.g. $H_2O$).

A mixture forms when two or more elements or compounds mix physically without any fixed ratio, and each component keeps its own chemical identity; mixtures can be separated by physical methods such as distillation, filtration, evaporation, crystallisation or magnetisation. A homogeneous mixture has uniform composition throughout (e.g. air, gasoline), while a heterogeneous mixture does not (e.g. soil, wood).

### 1.2.2 Atomic Number and Mass Number

The atomic number (Z) of an element equals the number of protons in the nucleus of its atoms; every atom of a given element has the same atomic number, which acts as its identification number (e.g. all hydrogen atoms have $Z = 1$, all carbon atoms have $Z = 6$). The mass number (A) is the sum of the number of protons and neutrons in the nucleus of an atom, calculated as $A = Z + n$, where n is the number of neutrons.

### 1.2.3 Relative Atomic Mass and the Atomic Mass Unit

Since the actual mass of an atom is too small to measure directly, atomic masses are compared against 1/12th the mass of a carbon-12 atom. This ratio is the relative atomic mass of an element, expressed in atomic mass units (amu), where 1 amu = $1.66 \times 10^{-24}$ g. On this scale, a proton has a mass of about 1.0073 amu, a neutron about 1.0087 amu, and an electron about $5.486 \times 10^{-4}$ amu.

### 1.2.4 Writing Chemical Formulae, Empirical and Molecular Formulae

A chemical formula is written by placing the positive ion first and the negative ion second, marking the valency of each ion at its top-right corner, then cross-exchanging those valencies to the lower-right corner of the other ion (e.g. $Na^+Cl^-$ becomes $NaCl$; $Ca^{2+}Cl^-$ becomes $CaCl_2$). If the valencies are equal they are dropped from the formula; if a radical is involved (like $SO_4^{2-}$), its net charge is used the same way, with the radical written inside parentheses when more than one is needed (e.g. $Al_2(SO_4)_3$).

The empirical formula is the simplest whole-number ratio of atoms (or ions, for an ionic compound) present in a compound -- for example, glucose has empirical formula $CH_2O$. The molecular formula shows the actual number of atoms of each element present in one molecule, and is related to the empirical formula by Molecular formula = (Empirical formula)n, where n is a whole number -- for example, benzene's molecular formula $C_6H_6$ comes from empirical formula CH with n = 6. Ionic compounds exist only as empirical formulae, called formula units (e.g. NaCl, KBr).

### 1.2.5 Molecular Mass and Formula Mass

Molecular mass is the sum of the atomic masses of all atoms in one molecule of a molecular substance (e.g. water, $H_2O$, has molecular mass 18 amu). Formula mass is used instead for ionic compounds, and is the sum of atomic masses of all atoms in one formula unit (e.g. sodium chloride has formula mass 58.5 amu, and $CaCO_3$ has formula mass 100 amu).

### 1.3.1 Ions, Molecular Ions and Free Radicals

An ion is an atom or group of atoms carrying a charge. A cation carries a positive charge, formed when an atom loses electrons from its outermost shell (e.g. $Na^+$, $K^+$); an anion carries a negative charge, formed when an atom gains electrons (e.g. $Cl^-$, $O^{2-}$). A molecular ion (or radical) forms when a molecule loses or gains an electron, giving it a net positive or negative charge (e.g. $CH_4^+$, $He^+$); cationic molecular ions are more common than anionic ones.

A free radical is an atom or group of atoms with an odd (unpaired) number of electrons, shown with a dot over the symbol (e.g. Cl(dot), $H_3C$(dot)); free radicals form by homolytic (equal) bond breakage when heat or light energy is absorbed, and are extremely reactive because they seek to complete their octet.

Molecules are classified by the number and type of atoms they contain: monoatomic (one atom, e.g. He, Ne, Ar), diatomic (two atoms, e.g. $H_2$, $O_2$, HCl), triatomic (three atoms, e.g. $H_2O$, $CO_2$), and polyatomic (many atoms, e.g. $CH_4$, $H_2SO_4$). A molecule made of the same type of atom is homoatomic (e.g. $O_3$, $S_8$); one made of different atoms is heteroatomic (e.g. $CO_2$, $NH_3$).

### 1.4 & 1.5 Avogadro's Number and the Mole

Avogadro's Number is a fixed collection of $6.02 \times 10^{23}$ particles, represented by the symbol $N_A$ -- named after the Italian scholar Amedeo Avogadro (1776-1856). $6.02 \times 10^{23}$ atoms of an element, $6.02 \times 10^{23}$ molecules of a molecular substance, or $6.02 \times 10^{23}$ formula units of an ionic compound are all equivalent to one mole of that substance, just as twelve eggs make one dozen.

A mole is defined as the amount of a substance that contains $6.02 \times 10^{23}$ particles, abbreviated 'mol'. When the atomic mass, molecular mass or formula mass of a substance is expressed in grams, that quantity is called its molar mass and equals one mole of the substance (e.g. 12 g of carbon = 1 mole of carbon; 18 g of $H_2O$ = 1 mole of water). The relationship between mass and moles is: Number of moles = known mass of substance / molar mass of substance.

### 1.6 Chemical Calculations

Two linked calculations recur throughout chemistry: mole-mass calculations (Number of moles = known mass / molar mass, and its rearrangement Mass = number of moles x molar mass) and mole-particle calculations (Number of moles = number of particles / $6.02 \times 10^{23}$, and its rearrangement Number of particles = number of moles x $6.02 \times 10^{23}$). To find the number of atoms in a molecular compound or the number of ions in an ionic compound, always calculate the number of moles first, then the number of molecules or formula units, and only then the number of atoms or ions -- never jump straight from mass to number of particles.

## Important Definitions

### What is chemistry?

The branch of science that deals with the composition, structure, properties and reactions of matter.

### Define matter.

Anything that has mass and occupies space; it can exist as a solid, liquid or gas.

### What is a substance?

A piece of matter in pure form, having a fixed composition and specific properties.

### What is a mixture?

Impure matter formed by combining two or more substances without a fixed ratio; it may be homogeneous or heterogeneous.

### Define an element.

A substance made up of the same type of atoms, having the same atomic number, that cannot be broken down into simpler substances by ordinary chemical means.

### Define a compound.

A substance made of two or more elements chemically combined together in a fixed ratio by mass, with properties entirely different from its constituent elements.

### What is valency?

The combining capacity of an element with other elements, depending on the number of electrons in its outermost (valence) shell.

### Define atomic number (Z).

The number of protons present in the nucleus of an atom of an element; it is the same for every atom of that element.

### Define mass number (A).

The sum of the number of protons and neutrons present in the nucleus of an atom, calculated as $A = Z + n$.

### What is relative atomic mass?

The average mass of the atoms of an element compared to 1/12th the mass of an atom of carbon-12.

### Define atomic mass unit (amu).

The unit for relative atomic mass, equal to 1/12th the mass of one atom of carbon-12; 1 amu = $1.66 \times 10^{-24}$ g.

### What is an empirical formula?

The simplest whole-number ratio of atoms (or ions) present in a compound.

### What is a molecular formula?

A formula that shows the actual number of atoms of each element present in one molecule of a compound.

### Define a formula unit.

The simplest whole-number ratio of ions present in an ionic compound, used to represent it since ionic compounds do not exist as independent molecules.

### Define a mole.

The amount of a substance that contains $6.02 \times 10^{23}$ particles (atoms, molecules or formula units); abbreviated 'mol'.

### What is Avogadro's Number?

A fixed collection of $6.02 \times 10^{23}$ particles, represented by the symbol $N_A$, equivalent to one mole of any substance.

### Define a cation and an anion.

A cation is an atom or group of atoms with a positive charge, formed by losing electrons. An anion is an atom or group of atoms with a negative charge, formed by gaining electrons.

### What is a free radical?

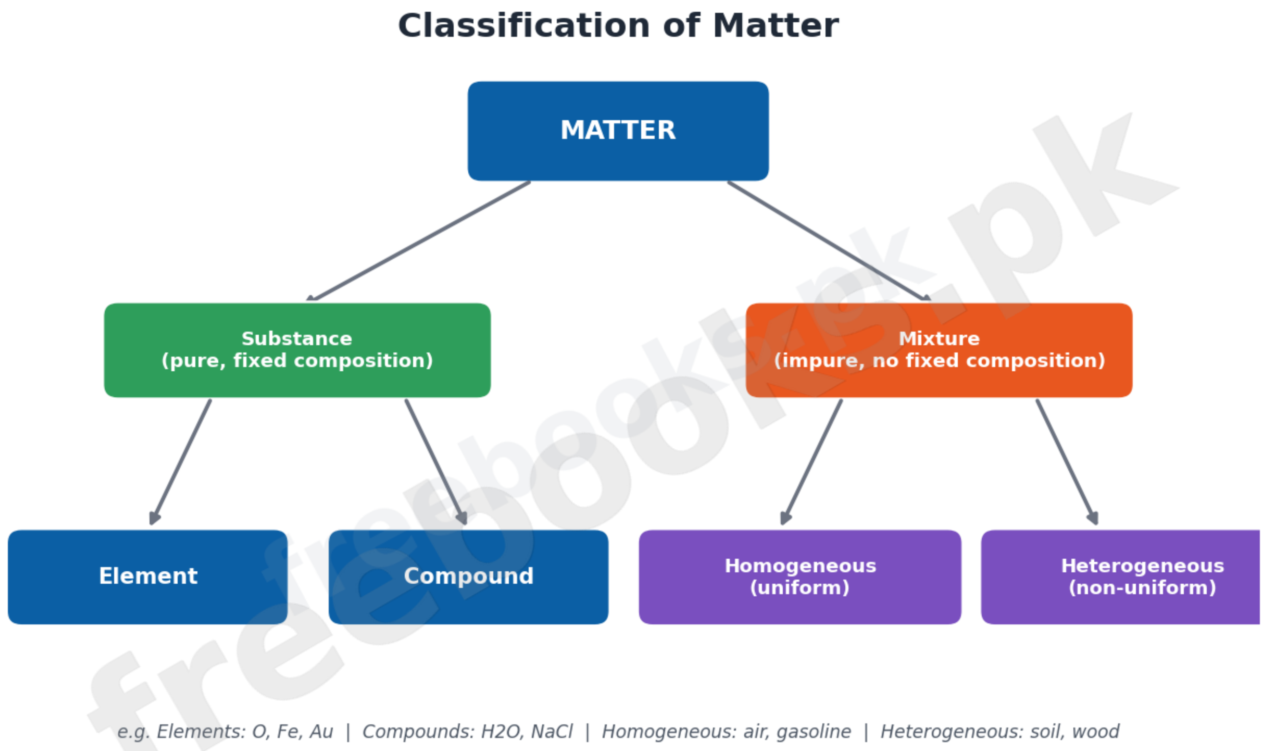
An atom or group of atoms possessing an odd (unpaired) number of electrons, represented by a dot over the element's symbol, and formed by homolytic bond breakage.

## Key Formulas

| Topic | Formula |
|---|---|
| Mass number | A = Z + n (n = number of neutrons) |
| Atomic mass unit | 1 amu = $1.66 \times 10^{-24}$ g |
| Molecular formula from empirical formula | Molecular formula = (Empirical formula)n |
| Number of moles from mass | Number of moles = known mass of substance / molar mass of substance |
| Mass from number of moles | Mass of substance (g) = number of moles x molar mass (g) |
| Number of moles from number of particles | Number of moles = given number of particles / $6.02 \times 10^{23}$ |
| Number of particles from number of moles | Number of particles = number of moles x $6.02 \times 10^{23}$ |
| Avogadro's Number | $N_A = 6.02 \times 10^{23}$ particles = 1 mole of any substance |

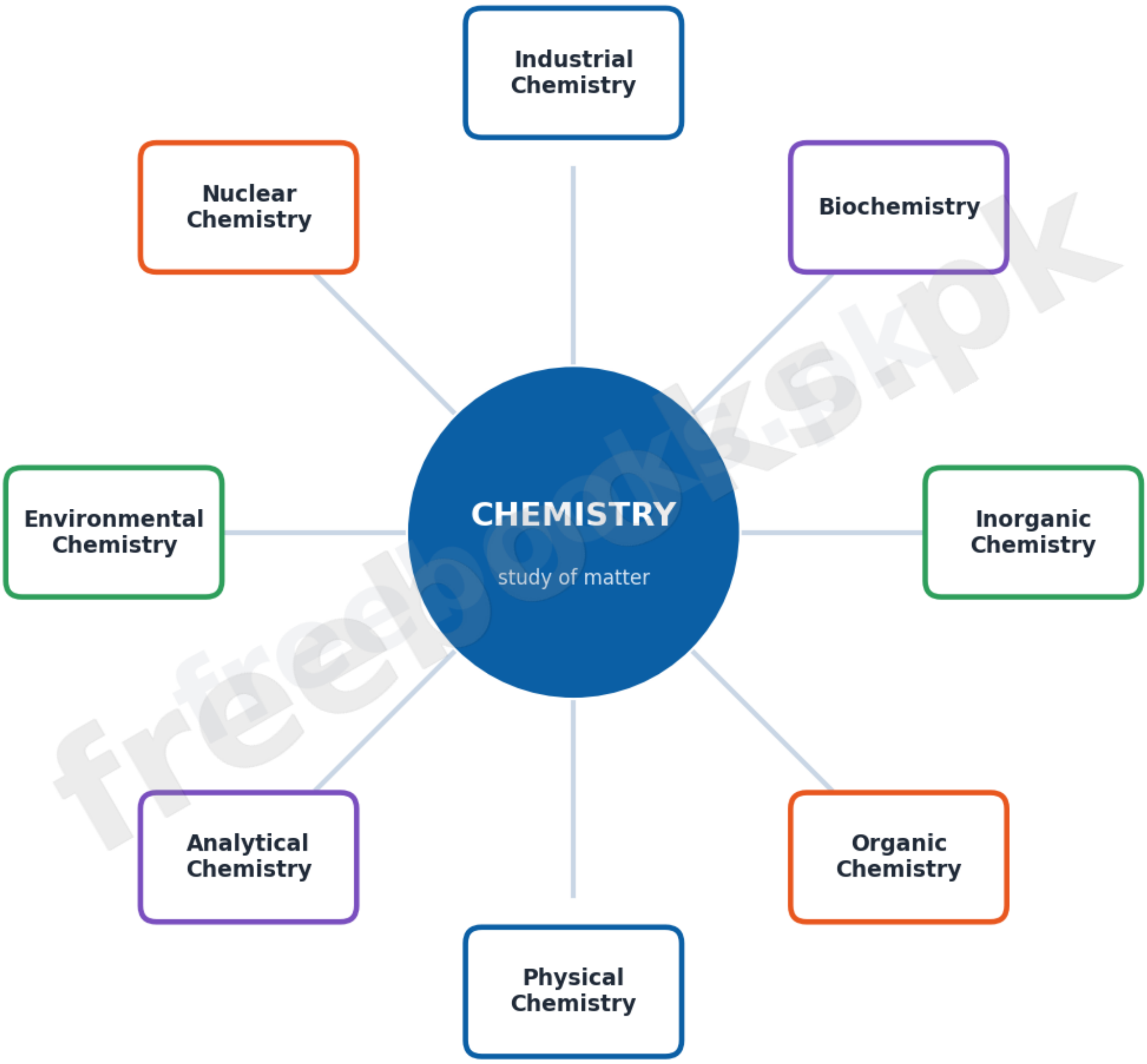
## Diagrams

**Classification of Matter:** Flowchart showing how matter divides into substances (elements and compounds) and mixtures (homogeneous and heterogeneous), with real-life examples of each.



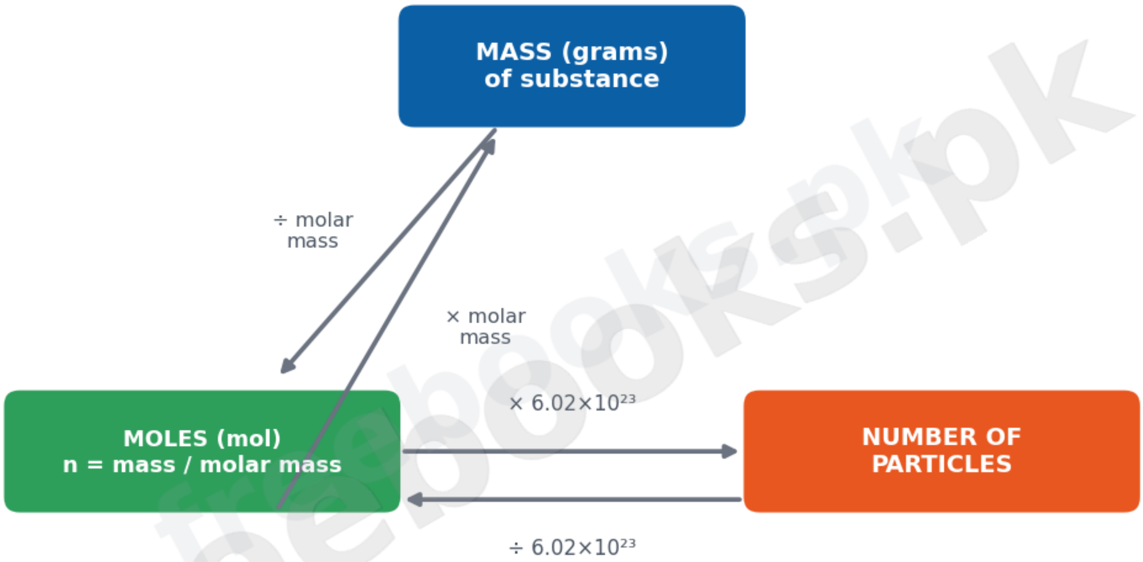

**The Eight Branches of Chemistry:** A radial diagram showing chemistry's eight main branches -- physical, organic, inorganic, biochemistry, industrial, nuclear, environmental and analytical chemistry -- around the central subject.

## The Eight Branches of Chemistry




**The Mole Concept: Mass, Moles and Particles:** A triangle diagram showing how to convert between mass (grams), number of moles, and number of particles using molar mass and Avogadro's Number (6.02 x 10^23).

## The Mole Concept: Mass ↔ Moles ↔ Particles




**1 mole = $6.02 \times 10^{23}$ particles (atoms, molecules or formula units) = Avogadro's Number (N_A)**

*particles include: atoms of an element, molecules of a molecular compound, or formula units of an ionic compound*

## Short Questions & Answers

### Which branch of chemistry studies the behaviour of gases, liquids and solids?

Physical chemistry studies the behaviour of gases, liquids and solids, along with the effect of temperature or radiation on matter.

### How does a compound differ from a mixture in terms of composition?

A compound has a fixed composition by mass and is represented by a chemical formula, while a mixture has no fixed composition and its components can be separated by simple physical methods.

### What is the difference between a homogeneous and a heterogeneous mixture?

A homogeneous mixture has uniform composition throughout, such as air or gasoline, while a heterogeneous mixture does not have uniform composition, such as soil or wood.

### Why is the atomic mass of an atom called its relative atomic mass?

Because the actual mass of an atom is too small to measure directly, so it is instead expressed relative to 1/12th the mass of a carbon-12 atom.

### Differentiate between an empirical formula and a molecular formula with one example.

An empirical formula gives the simplest whole-number ratio of atoms (e.g. $CH_2O$ for glucose), while a molecular formula gives the actual number of atoms in one molecule (e.g. $C_6H_6O_6$ for glucose).

### What is the difference between a cation and an anion?

A cation is a positively charged ion formed by losing electrons (e.g. $Na^+$), while an anion is a negatively charged ion formed by gaining electrons (e.g. $Cl^-$).

### How many particles are present in one mole of any substance?

One mole of any substance contains $6.02 \times 10^{23}$ particles (atoms, molecules or formula units), known as Avogadro's Number.

### Why can ionic compounds not be represented by a molecular formula?

Because ionic compounds do not exist as independent molecules; they form a three-dimensional crystal lattice and are represented instead by their simplest ratio, the formula unit.

### What is a free radical, and how is it represented?

A free radical is an atom or group of atoms with an odd (unpaired) number of electrons, represented by placing a dot over the element's symbol, e.g. Cl-dot.

### Distinguish between a molecule and a molecular ion.

A molecule is always neutral and is a stable unit formed by the combination of atoms, while a molecular ion carries a positive or negative charge and forms when a molecule loses or gains an electron, making it a reactive species.

## Long Questions & Answers

### Define element, compound and mixture, and explain how they are classified and distinguished from one another.

#### What is an element and how is it represented?

An element is a substance made up of the same type of atoms, having the same atomic number, that cannot be broken down into simpler substances by ordinary chemical means. Elements are represented by symbols derived from their English, Latin, Greek or German names -- a single capital letter (H, N, C) or a capital plus lowercase letter (Ca, Na, Cl). 118 elements are known today, of which 92 occur naturally, and about 80 percent are metals.

#### What is a compound and how does it form?

A compound is a substance formed when two or more elements chemically combine in a fixed ratio by mass. During this combination the elements lose their own individual properties and produce an entirely new substance -- for example, hydrogen and oxygen combine in a fixed 1:8 mass ratio to form water, which behaves nothing like either gas alone.

#### How is a mixture different from an element or compound?

A mixture forms when two or more elements or compounds are combined physically, without any fixed ratio, and each component retains its own chemical identity and properties. Unlike compounds, mixtures can be separated back into their original components using simple physical methods such as filtration, distillation, evaporation or crystallisation.

#### How are mixtures further classified?

Mixtures are classified as homogeneous, where the composition is uniform throughout (such as air, gasoline or ice cream), or heterogeneous, where the composition is not uniform throughout (such as soil, rock or wood).

## Explain the mole concept and describe how mass, number of moles, and number of particles of a substance are related.

### What is a mole?

A mole is the amount of a substance that contains $6.02 \times 10^{23}$ particles -- atoms, molecules, or formula units, depending on whether the substance is an element, a molecular compound, or an ionic compound. It is abbreviated 'mol' and provides a bridge between the microscopic world of particles and the macroscopic world of measurable mass.

### What is Avogadro's Number and why is it important?

Avogadro's Number, $6.02 \times 10^{23}$, represented by $N_A$, is the fixed number of particles present in exactly one mole of any substance. It plays the same role for chemists that 'a dozen' plays for counting eggs -- it allows an enormously large, fixed count of particles to be handled as a single convenient unit, the mole.

### How is the mass of a substance related to the number of moles?

The number of moles of a substance equals its known mass divided by its molar mass: Number of moles = known mass / molar mass. Rearranging this equation gives Mass of substance (g) = number of moles x molar mass (g), allowing conversion in either direction between a measured mass and a mole quantity.

### How is the number of moles related to the number of particles?

The number of moles of a substance equals the given number of particles divided by Avogadro's Number: Number of moles = number of particles / $6.02 \times 10^{23}$. Rearranging gives Number of particles = number of moles x $6.02 \times 10^{23}$. To find the number of atoms in a molecule or ions in a formula unit, the number of moles and then the number of molecules/formula units must be found first -- never jump directly from mass to number of particles.

## Multiple Choice Questions (MCQs)

**Industrial chemistry deals with the manufacturing of compounds: (A) in the laboratory (B) on a micro scale (C) on a commercial scale (D) on an economic scale**

Correct answer: (C) on a commercial scale. Industrial chemistry is defined as the branch that deals with the manufacturing of chemical compounds on a commercial scale.

**Which one of the following can be separated by physical means into its components? (A) a compound (B) an element (C) a mixture (D) a radical**

Correct answer: (C) a mixture. A mixture, unlike a compound or element, can be separated into its components by simple physical methods such as filtration or distillation.

**The most abundant element occurring in the oceans (by weight) is: (A) oxygen (B) hydrogen (C) nitrogen (D) silicon**

Correct answer: (A) oxygen. Oxygen makes up about 86% of the oceans by weight, as hydrogen and oxygen combine to form water.

**Which one of the following elements is found in greatest abundance in the Earth's crust? (A) oxygen (B) aluminium (C) silicon (D) iron**

Correct answer: (A) oxygen. Oxygen constitutes about 47% of the Earth's crust by weight, the highest of any single element.

**One atomic mass unit (amu) is equivalent to: (A) $1.66 \times 10^{-24}$ mg (B) $1.66 \times 10^{-24}$ g (C) $1.66 \times 10^{-24}$ kg (D) $1.66 \times 10^{-23}$ g**

Correct answer: (B) $1.66 \times 10^{-24}$ g. By definition, 1 amu = $1.66 \times 10^{-24}$ grams.

**Which one of the following molecules is not tri-atomic? (A) $H_2$ (B) $O_3$ (C) $H_2O$ (D) $CO_2$**

Correct answer: (A) $H_2$. $H_2$ is a diatomic molecule (two atoms), while $O_3$, $H_2O$ and $CO_2$ each contain three atoms.

**The molar mass of $H_2SO_4$ is: (A) 98 g (B) 98 amu (C) 9.8 g (D) 9.8 amu**

Correct answer: (A) 98 g. The molecular mass of $H_2SO_4$ is 98 amu, and when expressed in grams it becomes the molar mass, 98 g.

**An atom or group of atoms carrying a positive charge is called a/an: (A) anion (B) cation (C) free radical (D) isotope**

Correct answer: (B) cation. A cation is an atom or group of atoms carrying a positive charge, formed when electrons are lost from the outermost shell.

**The simplest whole-number ratio of atoms present in a compound is called its: (A) molecular formula (B) empirical formula (C) structural formula (D) formula unit only**

Correct answer: (B) empirical formula. The empirical formula gives the simplest whole-number ratio of atoms present in a compound.

**How many particles are contained in exactly one mole of a substance? (A) 6.02 x 10^22 (B) 6.02 x 10^23 (C) 1.66 x 10^-24 (D) 3.01 x 10^23**

Correct answer: (B) 6.02 x 10^23. One mole of any substance contains 6.02 x 10^23 particles, known as Avogadro's Number.

## Quick Revision Summary

- Chemistry has 8 branches: physical, organic, inorganic, biochemistry, industrial, nuclear, environmental, analytical
- Matter = substance (element or compound) OR mixture (homogeneous or heterogeneous)
- Atomic number (Z) = number of protons; Mass number (A) = protons + neutrons = Z + n
- 1 amu = 1.66 x 10^-24 g, based on 1/12th the mass of a carbon-12 atom
- Empirical formula = simplest whole-number ratio; Molecular formula = (Empirical formula)n
- Cation = positive ion (loses electrons); Anion = negative ion (gains electrons); Free radical = odd/unpaired electrons
- 1 mole = 6.02 x 10^23 particles (Avogadro's Number, N_A) = molar mass in grams
- Number of moles = mass / molar mass; Number of particles = moles x 6.02 x 10^23

## Exam Tips

- Memorise the eight branches of chemistry with one defining feature of each -- a common short-question topic
- Practice writing chemical formulae using the cross-exchange valency method until it becomes automatic
- Always calculate through moles first -- never jump directly from mass to number of particles or vice versa
- Be ready to classify given substances (element, compound or mixture; homogeneous or heterogeneous) -- a frequent exam question type
- Learn the standard values by heart: 1 amu = 1.66 x 10^-24 g and Avogadro's Number = 6.02 x 10^23
- Practice numericals on mole-mass and mole-particle calculations, since these carry heavy weightage in numerical questions

# Unit 2: Structure of Atoms

For most of the 19th century, Dalton's atomic theory held that atoms were indivisible, solid spheres. Experiments by Goldstein, J.J. Thomson, Rutherford and Chadwick between 1886 and 1932 overturned this idea, revealing that atoms are built from three subatomic particles -- electrons, protons and neutrons -- and are mostly empty space around a tiny, dense, positively charged nucleus.

This unit traces that discovery through the cathode-ray and canal-ray experiments, Rutherford's gold-foil experiment and planetary model, and Bohr's quantum-based model of fixed-energy orbits. It then covers how electrons are arranged into shells and subshells (electronic configuration) for the first 18 elements, and closes with isotopes -- atoms of the same element with different numbers of neutrons -- and their real-world uses in medicine, power generation, and dating fossils.

## Learning Objectives

- Describe the discovery of the electron, proton and neutron, and the experiments behind each
- Describe Rutherford's gold-foil experiment, his planetary atomic model, and its defects
- Explain how Bohr's atomic theory improved upon Rutherford's model using quantized energy levels
- Distinguish between shells and subshells, and state the maximum electron capacity of each
- Write the electronic configuration of the first 18 elements of the Periodic Table
- Define isotopes and compare the isotopes of hydrogen, carbon, chlorine and uranium
- State the importance and real-world uses of isotopes in medicine, power generation, and dating

## Key Concepts

### 2.1.1 Discovery of the Electron (Cathode Rays)

In 1897, J.J. Thomson identified negatively charged particles -- electrons -- while studying cathode rays produced in a discharge tube by Sir William Crookes in 1895. When high-voltage current passed through gas at very low pressure, shiny rays travelled from the cathode toward the anode. These cathode rays travel in straight lines, cast sharp shadows of opaque objects, deflect toward the positive plate in an electric field (proving they are negatively charged), raise the temperature of objects they strike, and are identical regardless of which gas or cathode material is used -- proving electrons are fundamental particles present in all matter. Thomson proposed the 'plum pudding' model: a solid sphere of positive charge with negative electrons embedded inside it, like plums in a pudding.

### 2.1.2 Discovery of the Proton (Canal Rays)

In 1886, Goldstein discovered a second type of ray in the discharge tube, travelling opposite to the cathode rays through holes in a perforated cathode -- he called these 'canal rays'. Canal rays deflect in electric and magnetic fields toward the negative plate (proving they are positively charged), and unlike cathode rays, their nature depends on the gas used in the tube. They form when cathode-ray electrons collide with and ionise residual gas molecules ($M + e^- \rightarrow M^+ + 2e^-$). The lightest gas, hydrogen, produces the lightest positive particles -- protons -- whose mass is about 1840 times that of an electron.

### 2.1.3 Discovery of the Neutron

Rutherford predicted in 1920 that atomic mass could not be explained by protons and electrons alone, so a neutral particle with a mass close to that of a proton must also exist. In 1932, Chadwick confirmed this by bombarding alpha particles onto a beryllium target ($^{9}Be + {}^{4}He \rightarrow {}^{12}C + {}^{1}n$), producing highly penetrating, uncharged radiation -- the neutron. Neutrons carry no charge, are highly penetrating, and have a mass nearly equal to that of a proton.

### 2.1.4 Rutherford's Atomic Model

Rutherford bombarded alpha particles (helium nuclei, $He^{2+}$) at an extremely thin (0.00004 cm) gold foil and observed the scattering pattern on a zinc-sulphide screen. Almost all particles passed straight through undeflected, but a few were deflected at large angles, and a very few bounced straight back. From this, Rutherford concluded: most of an atom's volume is empty space; there is a tiny, dense, positively charged centre called the nucleus; electrons revolve around the nucleus; and since an atom is neutral overall, the number of electrons equals the number of protons. Except for electrons, all fundamental particles within the nucleus are called nucleons.

Rutherford's model had two defects: classical theory predicted that revolving (charged) electrons should continuously emit energy and eventually spiral into the nucleus, meaning atoms should collapse -- but they do not; and continuously emitting electrons should produce a continuous spectrum, but atoms actually produce a line spectrum. These unresolved questions -- why atoms are stable, and why they give a line spectrum -- set the stage for Bohr's model.

### 2.1.5 Bohr's Atomic Theory

In 1913, Neils Bohr proposed a model built on Max Planck's Quantum Theory. Its key postulates: electrons revolve only in fixed circular orbits of specific ('quantized') energy, called energy levels; while in a given orbit, an electron neither emits nor absorbs energy; energy is absorbed when an electron jumps to a higher orbit and radiated when it jumps to a lower orbit, following Planck's equation $\Delta E = E2 - E1 = h\nu$ (h = Planck's constant, $6.63\times10^{-34}$ Js); and an electron can only occupy orbits where its angular momentum equals a whole-number multiple of $h/2\pi$: $mvr = nh/2\pi$, where n (1, 2, 3...) is the orbit or quantum number. Bohr's model explained both why atoms are stable (electrons don't lose energy while in an orbit) and why atoms produce a line spectrum (energy is only absorbed/released in fixed jumps).

### 2.2.1 Shells and Subshells

Electrons revolve around the nucleus in shells (energy levels), numbered by 'n' (1, 2, 3...) and labelled K, L, M, N. The K shell, closest to the nucleus, has the lowest energy; energy increases outward. Each shell contains one or more subshells (s, p, d, f), and within a shell, the s subshell fills before the p subshell. An s subshell holds a maximum of 2 electrons; a p subshell holds a maximum of 6 electrons. The maximum electron capacity of a shell is $2n^2$ -- so K (n=1) holds 2, L (n=2) holds 8, M (n=3) holds 18, and N (n=4) holds 32 electrons.

### 2.2.2 Electronic Configuration of the First 18 Elements

To write an electron configuration, three things must be known: the total number of electrons in the atom (equal to its atomic number), the order in which shells and subshells fill by increasing energy, and the maximum capacity of each shell/subshell. Electrons fill in the sequence 1s2, 2s2, 2p6, 3s2, 3p6... -- the number before each letter is the shell, the letter is the subshell, and the superscript is the number of electrons in it. For example, an atom with 11 electrons (sodium) fills K(2), L(8), leaving 1 electron for M, written simply as 2, 8, 1, or in full as 1s2 2s2 2p6 3s1. Ions adjust the electron count first: $Cl^-$ has 17+1=18 electrons (2, 8, 8), matching the noble gas argon.

### 2.3 Isotopes: Definition and Examples

Isotopes are atoms of the same element that have the same atomic number but different mass numbers -- meaning they have identical electronic configuration and proton count but a different number of neutrons. Because chemical properties depend on electronic configuration, isotopes of an element behave identically in chemical reactions, but because physical properties depend on mass, their physical properties (like density) differ.

Hydrogen has three isotopes: protium ($^{1}H$, 0 neutrons), deuterium ($^{2}H$ or D, 1 neutron), and tritium ($^{3}H$ or T, 2 neutrons) -- all with 1 proton and 1 electron. Carbon has three isotopes: $^{12}C$ (98.9% abundant, stable), $^{13}C$ (1.1% abundant, stable), and $^{14}C$ (radioactive, used in carbon dating) -- all with 6 protons. Chlorine has two isotopes, $^{35}Cl$ and $^{37}Cl$ (both with 17 protons). Uranium has three isotopes -- $^{234}U$, $^{235}U$, and $^{238}U$ (all with 92 protons) -- with $^{238}U$ making up nearly 99% of natural uranium.

### 2.3.3 Uses of Isotopes

Isotopes have wide-ranging real-world applications. In radiotherapy, P-32 and Sr-90 (which emit weakly penetrating beta radiation) treat skin cancer, while Co-60 (which emits strongly penetrating gamma rays) treats cancer deeper within the body. As tracers in medicine, Iodine-131 diagnoses goiter in the thyroid gland, and technetium monitors bone growth. In archaeology and geology, radioactive-isotope dating (such as radiocarbon dating using C-14) estimates the age of fossils and old carbon-containing objects. In chemical structure determination, C-14 labels CO2 to trace its path through photosynthesis into glucose. In power generation, controlled nuclear fission of U-235 (bombarded with slow neutrons, producing Barium-139, Krypton-94, and 3 neutrons plus energy) generates electricity in nuclear reactors.

## Important Definitions

**What is an atom, according to Dalton's original theory?**

A very small, indivisible particle that makes up all matter; atoms of the same element are alike and combine in different ways to form compounds.

**Define cathode rays.**

Negatively charged rays that travel from the cathode toward the anode in a discharge tube at low pressure; their study led to the discovery of the electron.

**Define canal rays.**

Positively charged rays that travel opposite to cathode rays through a perforated cathode in a discharge tube; their study led to the discovery of the proton.

**What is the nucleus of an atom?**

The tiny, dense, positively charged centre of an atom, discovered by Rutherford, where nearly all of the atom's mass is concentrated.

**Define a nucleon.**

Any fundamental particle found within the nucleus of an atom (protons and neutrons), excluding electrons.

**What is an energy level (shell)?**

A fixed orbit of specific, quantized energy in which an electron revolves around the nucleus without emitting or absorbing energy, according to Bohr's model.

### Define a subshell.

A division within a shell (labelled s, p, d, f) that holds a specific maximum number of electrons and has a slightly different energy from other subshells in the same shell.

### What is the maximum electron capacity of a shell?

$2n^2$, where n is the shell number -- so K holds 2, L holds 8, M holds 18, and N holds 32 electrons.

### Define electronic configuration.

The distribution of electrons of an atom among its shells and subshells, written in order of increasing energy, e.g. 1s2 2s2 2p6.

### Define isotopes.

Atoms of the same element having the same atomic number but different mass numbers, due to a different number of neutrons.

### What is Planck's constant?

A fundamental constant, h = 6.63 x 10^-34 Js, relating the energy of a photon to its frequency in Planck's equation, $\Delta E = h\nu$.

### Define quantum.

The smallest, fixed amount of energy that can be emitted or absorbed as electromagnetic radiation.

### What is radiocarbon (carbon) dating?

A method of estimating the age of old carbon-containing objects, such as fossils, by measuring the radioactivity of the isotope C-14 in them.

### Define nuclear fission.

A reaction in which a heavy atomic nucleus (such as U-235) splits into two lighter nuclei upon bombardment with a neutron, releasing a large amount of energy.

## Key Formulas

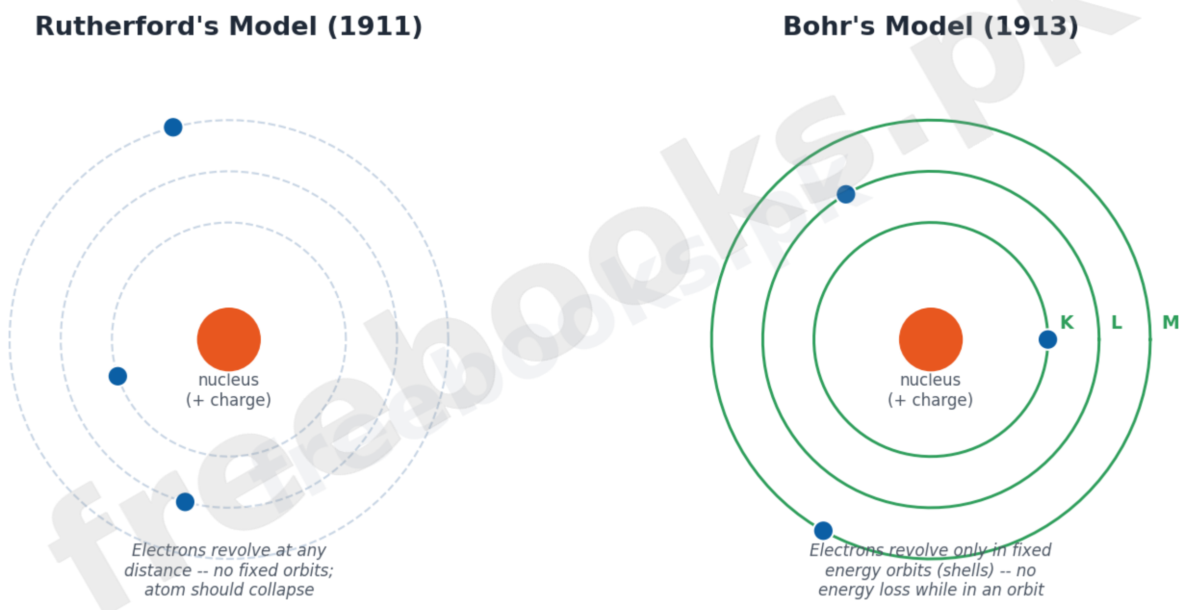
| Topic | Formula |
| --- | --- |
| Mass of proton relative to electron | Mass of a proton ≈ 1840 x mass of an electron |

| | |
|---|---|
| Ionisation producing canal rays | $M + e^- \rightarrow M^+ + 2e^-$ |
| Discovery of neutron (Chadwick) | 9-4Be + 4-2He → 12-6C + 1-0n |
| Planck's energy equation | ΔE = E2 - E1 = hν (h = 6.63 x 10^-34 Js) |
| Bohr's quantized angular momentum | mvr = nh / 2π (n = 1, 2, 3, ...) |
| Maximum electron capacity of a shell | Maximum electrons = $2n^2$ (K=2, L=8, M=18, N=32) |
| Nuclear fission of U-235 | 235-92U + 1-0n → 139-56Ba + 94-36Kr + 3(1-0n) + energy |

## Diagrams

**Rutherford's Model vs Bohr's Model:** Side-by-side comparison showing electrons at arbitrary distances in Rutherford's planetary model versus electrons confined to fixed-energy shells (K, L, M) in Bohr's quantized model.

**Rutherford's Model vs Bohr's Model of the Atom**




**Shells, Subshells and Maximum Electron Capacity:** Table-style diagram of shells K, L, M, N with their subshells and maximum electron capacity ($2n^2$).

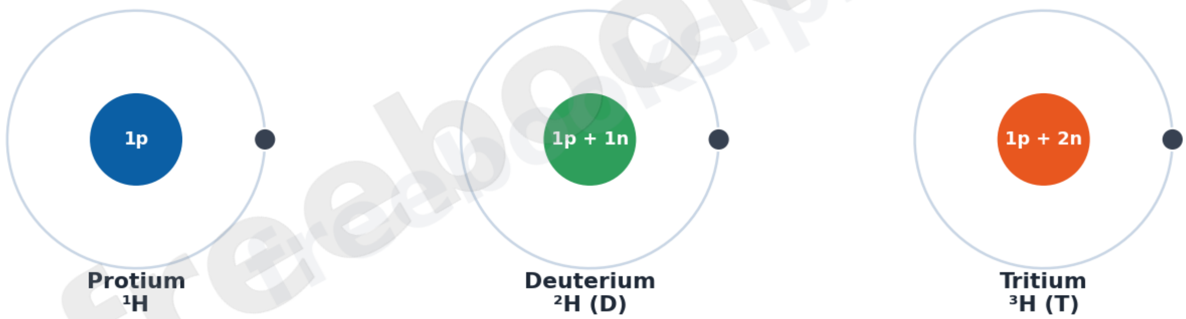
## Shells, Subshells and Maximum Electron Capacity

| Shell | n | Subshells present | Max. electrons ($2n^2$) |
|---|---|---|---|
| K | 1 | 1s | **2** |
| L | 2 | 2s, 2p | **8** |
| M | 3 | 3s, 3p, 3d | **18** |
| N | 4 | 4s, 4p, 4d, 4f | **32** |

*'s' subshell holds max. 2 electrons | 'p' subshell holds max. 6 electrons | filling order: s before p within the same shell*

**The Three Isotopes of Hydrogen:** Protium, deuterium and tritium shown with identical proton/electron counts but increasing numbers of neutrons.

### The Three Isotopes of Hydrogen (same protons, different neutrons)



## Short Questions & Answers

### What experiment led to the discovery of the electron, and who performed it?

J.J. Thomson's study of cathode rays produced in a discharge tube (originally built by Sir William Crookes) led to the discovery of the electron in 1897.

### Why are canal rays also called positive rays?

Because they deflect toward the negative plate in an electric field, showing they carry a positive charge, and they are formed by ionisation of residual gas molecules.

### What two observations led Rutherford to conclude that an atom is mostly empty space?

Almost all alpha particles passed straight through the gold foil undeflected, showing that most of an atom's volume must be empty, with only a tiny, dense nucleus present.

### Why does classical theory predict that atoms should collapse, but they do not?

Classical theory predicts that revolving charged electrons should continuously lose energy and spiral into the nucleus, but Bohr's model resolved this by showing electrons in a fixed orbit do not emit energy at all.

### What is the significance of Planck's equation, $\Delta E = h\nu$, in Bohr's model?

It shows that an electron only absorbs or radiates energy in fixed amounts (quanta) when it jumps between orbits, explaining why atoms produce a line spectrum rather than a continuous one.

### Why does the 's' subshell fill before the 'p' subshell in the same shell?

Because the 's' subshell has slightly lower energy than the 'p' subshell within the same shell, and electrons always occupy the lowest available energy first.

### What is the electronic configuration of an atom with 15 electrons?

2, 8, 5, since K holds 2, L holds 8, and the remaining 5 electrons go into the M shell (element: phosphorus).

### Why do isotopes of an element have identical chemical properties?

Because chemical properties depend on electronic configuration, and isotopes of the same element have the same number of electrons and the same configuration.

### Name the three isotopes of hydrogen and their neutron counts.

Protium (0 neutrons), deuterium (1 neutron), and tritium (2 neutrons) -- all with 1 proton and 1 electron.

### Give one medical use of a radioactive isotope.

Iodine-131 is used as a tracer to diagnose goiter in the thyroid gland; Co-60 is used in radiotherapy to treat cancer within the body.

## Long Questions & Answers

### Describe the discovery of the electron, proton and neutron, including the experiments and scientists responsible for each.

#### How was the electron discovered?

J.J. Thomson identified the electron in 1897 while studying cathode rays, produced by passing high-voltage current through gas at low pressure in a discharge tube built by Sir William Crookes. These rays travelled from the cathode toward the anode, deflected toward the positive plate in an electric field (proving a negative charge), and were identical regardless of the gas or cathode material used, showing electrons are a fundamental part of all matter.

#### How was the proton discovered?

In 1886, Goldstein observed a second set of rays in the discharge tube, travelling opposite to the cathode rays through holes in a perforated cathode, which he named canal rays. These rays deflected toward the negative plate (proving a positive charge), and their properties depended on the gas used -- the lightest gas, hydrogen, produced the lightest positive particles, called protons.

#### How was the neutron discovered?

Rutherford predicted in 1920 that a neutral particle must exist in the atom to account for atomic mass. In 1932, Chadwick confirmed this by bombarding alpha particles onto a beryllium target, producing highly penetrating, uncharged radiation that he identified as the neutron, with a mass nearly equal to that of a proton.

#### What do these three discoveries establish together?

Together, they proved that Dalton's idea of an indivisible atom was incomplete -- an atom is actually built from three subatomic particles: negatively charged electrons, positively charged protons, and neutral neutrons, laying the foundation for Rutherford's and Bohr's atomic models.

### Explain Rutherford's atomic model and how Bohr's atomic theory resolved its defects.

#### What did Rutherford's gold-foil experiment show?

Rutherford bombarded alpha particles at an extremely thin gold foil. Since almost all particles passed straight through, he concluded that most of an atom's volume is empty space; since a few particles deflected at large angles or bounced back, he concluded that there is a tiny, dense, positively charged nucleus at the atom's centre, around which electrons revolve.

#### What were the defects of Rutherford's model?

According to classical theory, a revolving charged electron should continuously emit energy and eventually spiral into the nucleus, meaning atoms should collapse -- but they clearly do not. Also, continuously emitting electrons should produce a continuous spectrum, but atoms are observed to produce a line spectrum instead.

#### What did Bohr propose to fix these defects?

In 1913, Bohr proposed that electrons revolve only in fixed, quantized orbits of specific energy (energy levels), and that as long as an electron stays in one orbit, it neither absorbs nor emits energy -- resolving the collapse problem. Energy is only absorbed or released in a fixed amount, $\Delta E = h\nu$, when an electron jumps between orbits, explaining the line spectrum.

**What rule governs which orbits an electron can occupy?**

Bohr proposed that an electron can only occupy orbits where its angular momentum is a whole-number multiple of $h/2\pi$, expressed as $mvr = nh/2\pi$, where n is the orbit or quantum number (1, 2, 3...). This quantization rule explains why only certain fixed orbits, and not any arbitrary distance from the nucleus, are allowed.

## Multiple Choice Questions (MCQs)

**Which one of the following results in the discovery of the proton? (A) cathode rays (B) canal rays (C) X-rays (D) alpha rays**

Correct answer: (B) canal rays. Canal rays, discovered by Goldstein in 1886, were found to be positively charged and led to the discovery of the proton.

**Which one of the following is the most penetrating? (A) protons (B) electrons (C) neutrons (D) alpha particles**

Correct answer: (C) neutrons. Neutrons, being neutral and highly penetrating, pass through matter more easily than charged protons, electrons or alpha particles.

**The concept of fixed-energy orbits was introduced by: (A) J.J. Thomson (B) Rutherford (C) Bohr (D) Planck**

Correct answer: (C) Bohr. Neils Bohr introduced the concept of fixed, quantized energy orbits in his 1913 atomic model.

**Which one of the following shells consists of three subshells (s, p, d)? (A) K shell (B) L shell (C) M shell (D) N shell only up to p**

Correct answer: (C) M shell. The M shell (n=3) contains three subshells: 3s, 3p and 3d.

**Which radioisotope is used for the diagnosis of a tumor/goiter in the body? (A) cobalt-60 (B) iodine-131 (C) strontium-90 (D) phosphorus-32**

Correct answer: (B) iodine-131. Iodine-131 is used as a tracer to diagnose goiter in the thyroid gland.

**When U-235 undergoes fission, it produces: (A) only electrons (B) neutrons, along with Ba-139 and Kr-94 (C) only protons (D) no particles**

Correct answer: (B) neutrons, along with Ba-139 and Kr-94. The fission of U-235 with a slow neutron produces Barium-139, Krypton-94, three neutrons, and a large amount of energy.

### The p subshell can hold a maximum of: (A) 2 electrons (B) 6 electrons (C) 8 electrons (D) 10 electrons

Correct answer: (B) 6 electrons. A p subshell can accommodate a maximum of 6 electrons.

### Deuterium is used to make: (A) light water (B) heavy water (C) soft water (D) hard water

Correct answer: (B) heavy water. Deuterium ($^2H$) combines with oxygen to form heavy water (D2O).

### The isotope C-12 is present in nature in an abundance of approximately: (A) 96.9% (B) 97.6% (C) 98.9% (D) 99.9%

Correct answer: (C) 98.9%. Carbon-12 makes up about 98.9% of naturally occurring carbon.

### Who is credited with discovering the proton? (A) Goldstein (B) J.J. Thomson (C) Neils Bohr (D) Rutherford

Correct answer: (A) Goldstein. Goldstein discovered canal rays in 1886, which were later shown to be made of protons.

## Quick Revision Summary

- Cathode rays (Crookes/Thomson) -> discovery of the electron (negative charge)
- Canal rays (Goldstein) -> discovery of the proton (positive charge, depends on gas used)
- Chadwick (1932): alpha particles + beryllium -> discovery of the neutron (no charge)
- Rutherford's gold-foil experiment: mostly empty space + tiny dense positive nucleus
- Rutherford's defects: atom should collapse; should show continuous spectrum (it doesn't)
- Bohr (1913): electrons revolve in fixed, quantized-energy orbits; $\Delta E = h\nu$ between orbits

- Shell capacity = $2n^2$ -> K=2, L=8, M=18, N=32; 's' subshell fills before 'p' in the same shell
- Isotopes = same atomic number, different mass number (different neutrons); same chemical, different physical properties
- H has 3 isotopes (protium, deuterium, tritium); C has 3 (12C, 13C, 14C); Cl has 2; U has 3

## Exam Tips

- Memorise which scientist is linked to which discovery/model: Goldstein (proton), Thomson (electron), Rutherford (nucleus), Bohr (energy levels), Chadwick (neutron)
- Practice writing electronic configurations for atomic numbers 1-18 and for common ions ($Na^+$, $Cl^-$, $Al^{3+}$, $Mg^{2+}$)
- Be ready to state at least 2 defects of Rutherford's model and how Bohr's model resolved each one
- Learn the shell capacity formula $2n^2$ and be able to apply it to K, L, M, N shells
- Practice the isotopes table for H, C, Cl and U (atomic number, mass number, protons, neutrons) -- a common short-question topic
- Know at least one real-world use of isotopes in each field: medicine, radiotherapy, dating, and power generation

# Unit 3: Periodic Table and Periodicity of Properties

Nineteenth-century chemists searched for a systematic way to organise the known elements. Early attempts -- Dobereiner's triads, Newlands' law of octaves, and finally Mendeleev's periodic table based on atomic mass -- gradually converged on the periodic law. In 1913, Moseley's discovery of atomic number replaced atomic mass as the true basis for arranging elements, giving us the modern periodic table.

This unit covers the structure of the modern (long form) periodic table -- its periods, groups, and s/p/d/f blocks -- and then examines periodicity: how atomic size, shielding effect, ionization energy, electron affinity, and electronegativity change in predictable trends across a period and down a group.

## Learning Objectives

- Distinguish between a period and a group in the Periodic Table
- State the periodic law in both its original (Mendeleev) and modern (Moseley) forms
- Classify elements into periods and groups according to their outermost electron configuration
- Determine the demarcation of the periodic table into s-block, p-block, d-block and f-block
- Explain the shape and salient features of the long form periodic table
- Recognise the similarity in physical and chemical properties of elements in the same family
- Explain how the shielding effect influences periodic trends
- Describe how atomic radius, ionization energy, electron affinity and electronegativity change across a period and down a group

## Key Concepts

### 3.1.1 Early Attempts: Dobereiner, Newlands, Mendeleev

Dobereiner grouped elements into triads of three, where the middle element's atomic mass was the average of the other two (e.g. calcium 40, strontium 88, barium 137) -- but only a few elements fit this pattern. Newlands' 1864 'law of octaves' noted that chemical properties repeated every eighth element when arranged by increasing atomic mass, but left no room for undiscovered elements or the not-yet-known noble gases.

Mendeleev arranged the 63 known elements in order of increasing atomic mass into horizontal periods, placing elements of similar properties in the same vertical group, and stated the periodic law as 'properties of the elements are periodic functions of their atomic masses'. His table's major achievement was predicting the properties of undiscovered elements, though it had demerits: it could not explain the position of isotopes, and the atomic-mass ordering placed a few elements out of their correct order.

### 3.1.2 Periodic Law and the Modern Periodic Table

In 1913, H. Moseley discovered atomic number and showed it -- not atomic mass -- should determine an element's position, amending the periodic law to: 'properties of the elements are periodic functions of their atomic numbers'. Atomic number is more fundamental than atomic mass because it is fixed for every element and increases regularly by exactly 1 from element to element; no two elements share an atomic number.

The modern periodic table arranges elements by increasing atomic number. Properties repeat at regular intervals -- every eighth element initially (e.g. sodium, Z=11, resembles lithium, Z=3), and every eighteenth element after Z=18 -- so the long row of elements is cut into rows and stacked to form a table of periods (horizontal rows) and groups (vertical columns, numbered 1 to 18).

### 3.1.3 Periods, Groups and Blocks

A period's elements have continuously increasing atomic number and continuously changing electronic configuration, so their properties change steadily across the period; the number of valence electrons decides an element's position (1 valence electron = leftmost/alkali metals, 8 valence electrons = rightmost/noble gases). A group's elements do NOT have continuously increasing atomic numbers, but they do share the same number of valence electrons and therefore similar chemical properties -- which is why a group is also called a family.

The long form periodic table has 7 periods and 18 groups. Period 1 is a short period (2 elements: H, He). Periods 2 and 3 are normal periods (8 elements each). Periods 4 and 5 are long periods (18 elements each). Periods 6 and 7 are very long periods (32 elements each, though period 7 remains incomplete); their 14-element lanthanide and actinide series are placed separately below the main table to keep it a manageable size.

Elements are also classified into four blocks by which subshell receives the last electron: s-block (groups 1-2, valence electrons in an s subshell), p-block (groups 13-18, valence electrons in a p subshell), d-block (groups 3-12, transition elements, filling a d subshell), and f-block (lanthanides and actinides, filling an f subshell).

### 3.2.1 Atomic Size (Atomic Radius)

Since atoms have no fixed boundary, atomic radius is defined as half the distance between the nuclei of two bonded atoms of the same element (e.g. two carbon atoms are 154 pm apart, so carbon's atomic radius is 77 pm). Moving left to right across a period, atomic size decreases (e.g. Li 152 pm down to Ne 69 pm) because each added proton increases the effective nuclear charge while electrons keep filling the same shell, pulling the shell in tighter. Moving down a group, atomic size increases because a new electron shell is added at each period, and this greater distance and shielding outweighs the increase in nuclear charge.

### 3.2.2 Shielding Effect

Inner-shell electrons partially block ('shield') the nucleus's pull on the outermost (valence) electrons, so a valence electron feels a reduced, effective nuclear charge (Zeff) rather than the full nuclear charge. This is called the shielding effect. It increases down a group as more electrons and shells are added (e.g.

potassium, Z=19, has a stronger shielding effect than sodium, Z=11, so it is easier to remove an electron from potassium). Shielding effect decreases across a period from left to right, since electrons are added to the same shell rather than a new one.

### 3.2.3 Ionization Energy

Ionization energy is the energy required to remove the most loosely bound electron from an isolated gaseous atom (e.g. $Na \rightarrow Na^+ + e^-$, $\Delta H = +496$ kJ/mol for sodium's first ionization energy). Removing further electrons requires progressively more energy (second, third ionization energy, etc.). Ionization energy increases across a period left to right (e.g. Li 520 to Ne 2081 kJ/mol) because atoms get smaller and valence electrons are held more tightly by the increasing nuclear charge. It decreases down a group (e.g. Li 520 down to Cs 377 kJ/mol) because additional inner shells increase shielding and distance, making valence electrons easier to remove.

### 3.2.4 Electron Affinity

Electron affinity is the energy released when an electron is added to the outermost shell of an isolated gaseous atom (e.g. $F + e^- \rightarrow F^-$, $\Delta H = -328$ kJ/mol). It increases (becomes more negative/more energy released) across a period left to right, because decreasing atomic size means the nucleus attracts the incoming electron more strongly. It decreases down a group, because increasing atomic size and shielding effect weaken the nucleus's pull on the incoming electron (e.g. iodine's electron affinity is smaller in magnitude than chlorine's).

### 3.2.5 Electronegativity

Electronegativity is an atom's ability to attract the shared pair of electrons toward itself within a covalent bond -- important when analysing covalent bonding. Its trend matches ionization energy and electron affinity: electronegativity increases across a period left to right (e.g. Li 1.0 up to F 4.0) because a higher effective nuclear charge pulls the shared electron pair closer, and it decreases down a group (e.g. F 4.0 down to I 2.7) because increasing atomic size weakens the atom's pull on the shared electrons.

## Important Definitions

### State the (modern) periodic law.

The properties of the elements are periodic functions of their atomic numbers.

### What is a period in the Periodic Table?

A horizontal row of elements with continuously increasing atomic number and continuously changing electronic configuration and properties.

### What is a group in the Periodic Table?

A vertical column of elements that share the same number of valence electrons and therefore similar chemical properties; also called a family.

### What is a block of elements?

A set of elements classified by which type of subshell (s, p, d or f) receives the last electron in their electronic configuration.

### Define atomic radius.

Half of the distance between the nuclei of two bonded atoms of the same element.

### What is effective nuclear charge (Zeff)?

The actual nuclear charge felt by a valence electron after accounting for the shielding effect of inner-shell electrons.

### Define the shielding effect.

The reduction in the nuclear charge felt by valence electrons due to the presence of electrons in the inner shells between them and the nucleus.

### Define ionization energy.

The amount of energy required to remove the most loosely bound electron from the valence shell of an isolated gaseous atom.

### Define electron affinity.

The amount of energy released when an electron is added to the outermost shell of an isolated gaseous atom.

### Define electronegativity.

The ability of an atom to attract the shared pair of electrons toward itself within a molecule.

### What are transition elements?

Elements of groups 3 to 12, in which the d subshell is in the process of being filled with electrons.

### What are the lanthanides and actinides?

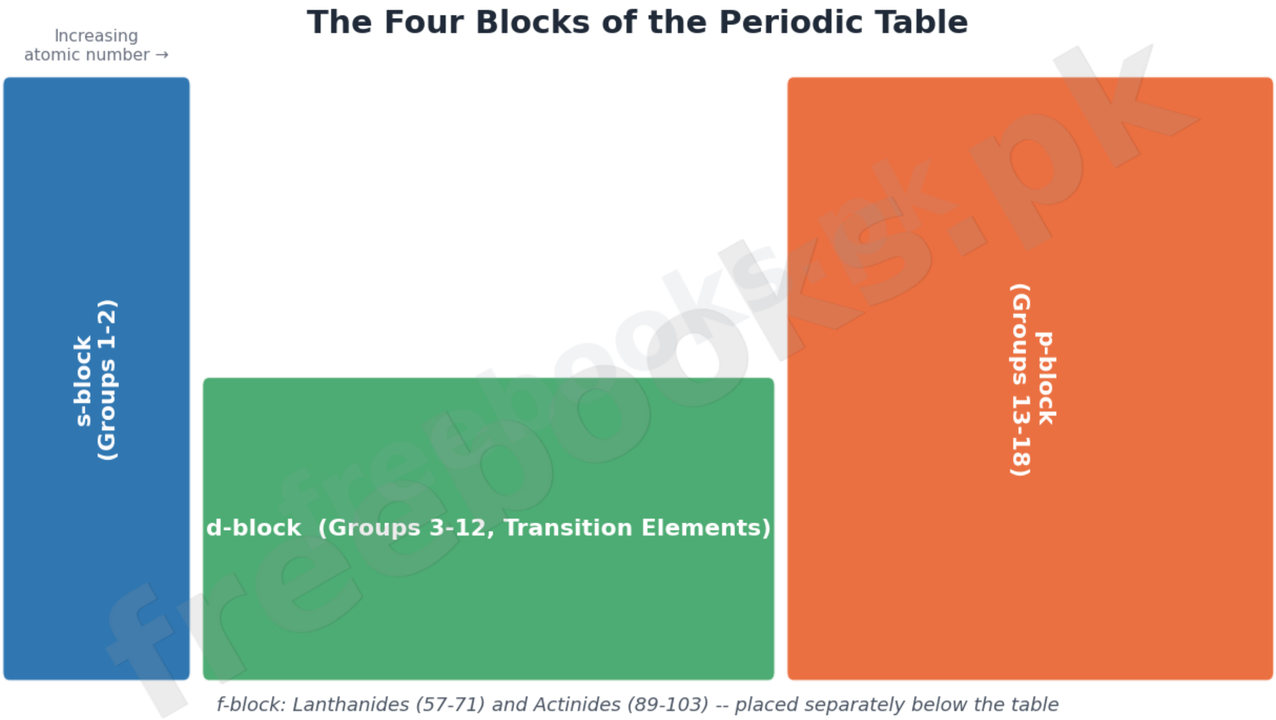
Two series of 14 elements each, starting after Lanthanum (Z=57) and Actinium (Z=89), placed separately below the main periodic table and together forming the f-block.

## Key Formulas

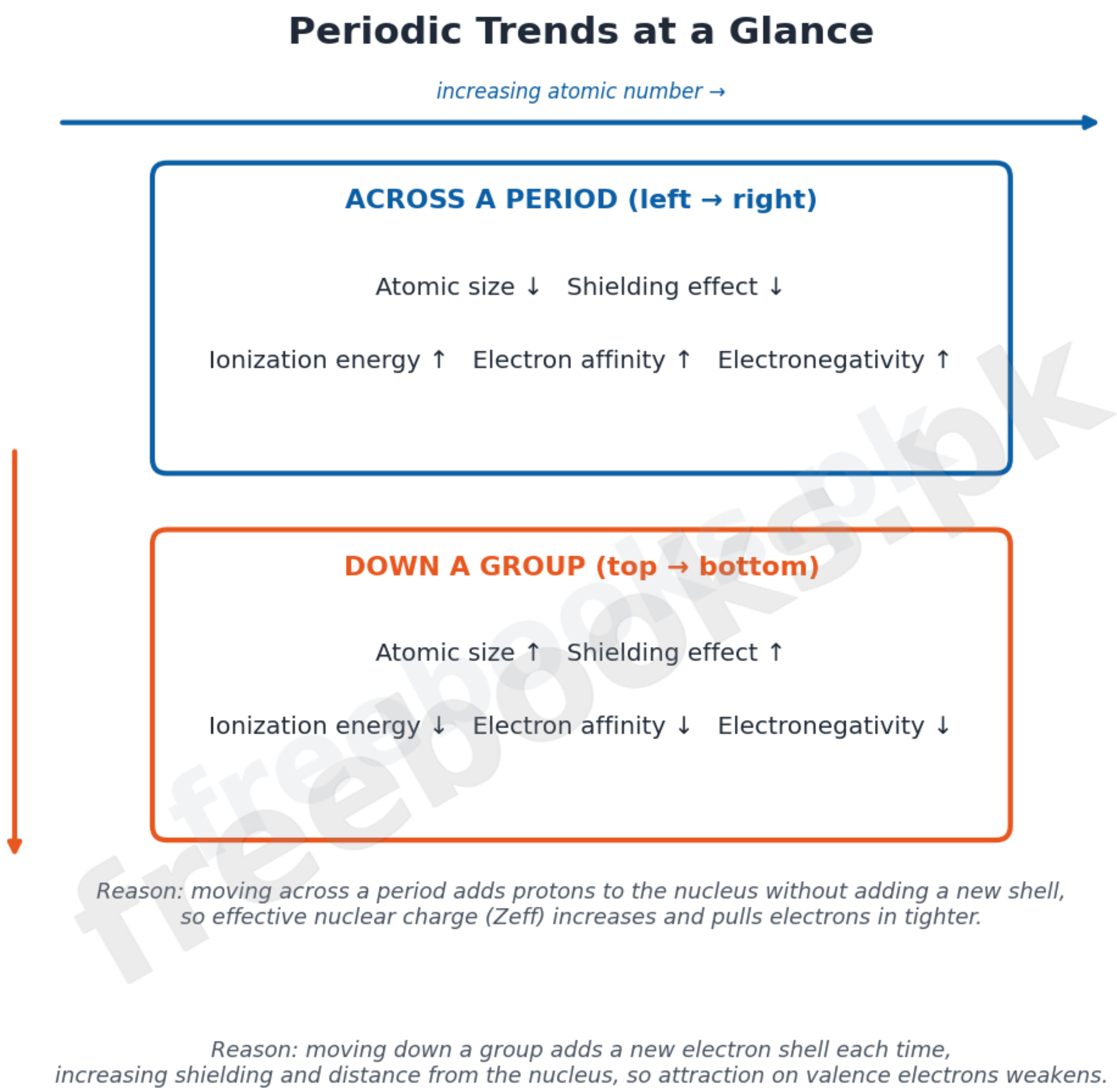
| Topic | Formula |
|---|---|
| Ionization energy example (sodium) | $Na \rightarrow Na^+ + e^-$, $\Delta H = +496$ kJ $mol^{-1}$ |
| Electron affinity example (fluorine) | $F + e^- \rightarrow F^-$, $\Delta H = -328$ kJ $mol^{-1}$ |
| Atomic size trend | Decreases across a period (left→right); increases down a group |
| Ionization energy trend | Increases across a period (left→right); decreases down a group |
| Electron affinity trend | Increases (more negative) across a period; decreases down a group |
| Electronegativity trend | Increases across a period (left→right); decreases down a group |
| Number of periods / groups | 7 periods, 18 groups in the modern long-form Periodic Table |

## Diagrams

**The Four Blocks of the Periodic Table:** s-block, p-block, d-block and f-block shown by which subshell type receives the last electron in each element group.



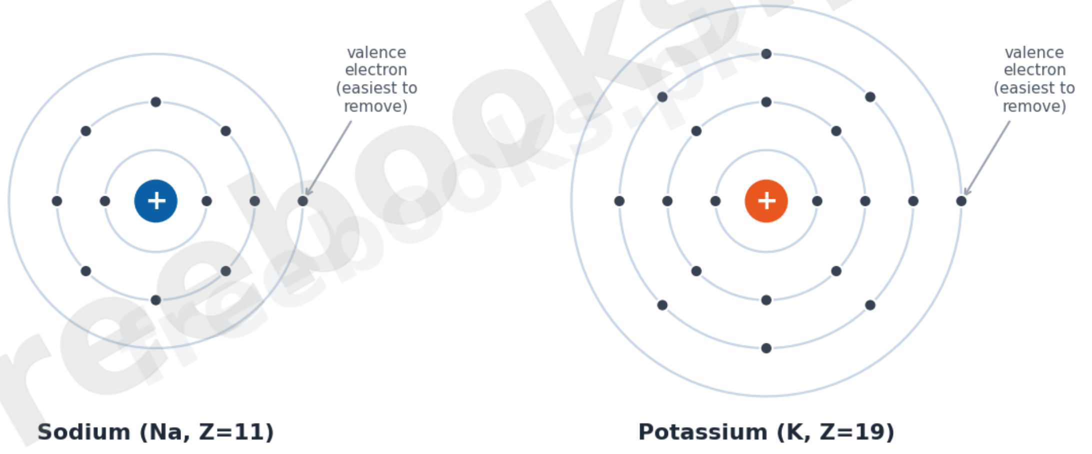

**Periodic Trends at a Glance:** Summary of how atomic size, shielding, ionization energy, electron affinity and electronegativity change across a period and down a group.






**Shielding Effect: Sodium vs Potassium:** Electron-shell diagrams comparing sodium and potassium, showing how potassium's extra inner shell increases shielding and makes its valence electron easier to remove.

Shielding Effect: More Inner Shells in K Weaken the Nucleus's Pull on the Outer Electron




## Short Questions & Answers

### What was the main achievement of Mendeleev's periodic table, despite its demerits?

It successfully predicted the properties of elements that had not yet been discovered at the time.

### Why is atomic number considered more fundamental than atomic mass for arranging elements?

Because atomic number is fixed for every element and increases regularly by exactly 1 from element to element, while no two elements can share the same atomic number.

### How many elements are in period 1, and why is it called a short period?

Period 1 has only 2 elements, hydrogen and helium, because the first shell (K) can hold a maximum of only 2 electrons.

### Why do elements in the same group have similar chemical properties?

Because elements in the same group have the same number of valence electrons, even though their atomic numbers increase with irregular gaps.

### Which block do group 1 and group 2 elements belong to, and why?

They belong to the s-block, because their valence electrons occupy an s subshell.

**Why does atomic size decrease across a period from left to right?**
Because the effective nuclear charge increases as protons are added to the nucleus, while electrons keep filling the same shell, pulling the shell inward.

**Why is potassium's valence electron easier to remove than sodium's?**
Because potassium has an additional inner shell compared to sodium, giving it a greater shielding effect that weakens the nucleus's pull on its outermost electron.

**Why does ionization energy decrease down a group?**
Because additional electron shells increase both the atomic size and the shielding effect, so the valence electron is held less tightly and is easier to remove.

**Why does electron affinity increase across a period from left to right?**
Because atomic size decreases across a period, so the nucleus attracts an incoming electron more strongly, releasing more energy.

**Which element in period 2 has the highest electronegativity, and why?**
Fluorine has the highest electronegativity in period 2 (4.0), because it has the smallest atomic size and highest effective nuclear charge among period 2 elements, giving it the strongest pull on a shared electron pair.

## Long Questions & Answers

### Trace the historical development of the periodic table from Dobereiner to the modern periodic law.

#### What did Dobereiner and Newlands contribute?

Dobereiner grouped elements into triads of three, where the middle element's atomic mass was the average of the outer two, but this pattern held for only a few elements. Newlands proposed the 'law of octaves' in 1864, noting that chemical properties repeated every eighth element when ordered by increasing atomic mass, comparing it to musical notes -- but his scheme left no room for undiscovered elements or the noble gases.

### What did Mendeleev achieve, and what were his table's demerits?

Mendeleev arranged the 63 known elements by increasing atomic mass into periods and groups, stating the periodic law as 'properties of elements are periodic functions of their atomic masses'. His table's major strength was predicting the properties of undiscovered elements. Its demerits were that it could not explain the position of isotopes, and a few elements were placed out of correct order when strictly following atomic mass.

### What change did Moseley bring, and why was it significant?

In 1913, Moseley discovered atomic number and showed that it, not atomic mass, should determine an element's position in the table. This amended the periodic law to 'properties of elements are periodic functions of their atomic numbers' -- atomic number is more fundamental because it is fixed and increases by exactly 1 for each successive element, with no two elements sharing a value.

### How did this lead to the modern periodic table?

Arranging elements by increasing atomic number showed that properties repeat at regular intervals (electronic configuration is directly tied to atomic number), so long rows of elements could be cut and stacked into periods and groups, producing the modern long-form periodic table with its s, p, d and f blocks.

## Explain how atomic size, ionization energy and electronegativity vary across a period and down a group, and why.

### How and why does atomic size change?

Atomic size decreases across a period from left to right, because each added proton increases the effective nuclear charge while electrons continue filling the same shell, pulling the shell inward (e.g. Li 152 pm down to Ne 69 pm). It increases down a group because a new electron shell is added at each successive period, increasing both distance from the nucleus and shielding effect.

### How and why does ionization energy change?

Ionization energy increases across a period from left to right, because decreasing atomic size means valence electrons are held more tightly by a stronger effective nuclear charge (e.g. Li 520 up to Ne 2081 kJ/mol). It decreases down a group, because additional inner shells increase shielding and distance from the nucleus, making the valence electron easier to remove (e.g. Li 520 down to Cs 377 kJ/mol).

### How and why does electronegativity change?

Electronegativity increases across a period from left to right, following the same reasoning as ionization energy: a higher effective nuclear charge pulls a shared pair of electrons more strongly toward the atom (e.g. Li 1.0 up to F 4.0). It decreases down a group because increasing atomic size places the shared electron pair farther from the nucleus, weakening the atom's pull on it (e.g. F 4.0 down to I 2.7).

### What single underlying factor connects all three trends?

All three trends are governed by the same balance between effective nuclear charge and atomic size/shielding: across a period, nuclear charge increases while shielding stays roughly constant, strengthening the nucleus's pull; down a group, an added shell increases both size and shielding faster than nuclear charge grows, weakening the nucleus's pull on outer electrons.

## Multiple Choice Questions (MCQs)

### The atomic radii of the elements in the Periodic Table: (A) increase from left to right in a period (B) increase from top to bottom in a group (C) do not change from left to right in a period (D) decrease from top to bottom in a group

Correct answer: (B) increase from top to bottom in a group. Atomic radii increase from top to bottom in a group because a new electron shell is added at each period.

**The amount of energy given out when an electron is added to an atom is called: (A) lattice energy (B) ionization energy (C) electronegativity (D) electron affinity**

Correct answer: (D) electron affinity. Electron affinity is defined as the energy released when an electron is added to an isolated gaseous atom.

**Mendeleev's Periodic Table was based upon the: (A) electronic configuration (B) atomic mass (C) atomic number (D) completion of a subshell**

Correct answer: (B) atomic mass. Mendeleev arranged elements in order of increasing atomic mass.

**The long form of the Periodic Table is constructed on the basis of: (A) Mendeleev's postulate (B) atomic number (C) atomic mass (D) mass number**

Correct answer: (B) atomic number. The modern long form periodic table is arranged according to increasing atomic number.

**The 4th and 5th periods of the long form Periodic Table are called: (A) short periods (B) normal periods (C) long periods (D) very long periods**

Correct answer: (C) long periods. Periods 4 and 5 are called long periods, each containing 18 elements.

**Which one of the following halogens has the lowest electronegativity? (A) fluorine (B) chlorine (C) bromine (D) iodine**

Correct answer: (D) iodine. Electronegativity decreases down group 17, so iodine has the lowest electronegativity among these halogens.

**Along a period (left to right), which one of the following decreases? (A) atomic radius (B) ionization energy (C) electron affinity (D) electronegativity**

Correct answer: (A) atomic radius. Atomic radius decreases across a period, while ionization energy, electron affinity and electronegativity all increase.

**Transition elements are: (A) all gases (B) all metals (C) all non-metals (D) all metalloids**

Correct answer: (B) all metals. All transition elements (groups 3-12, d-block) are metals.

**Which of the following statements about ionization energy is incorrect? (A) it is measured in kJ/mol (B) it involves absorption of energy (C) it decreases across a period (D) it decreases down a group**

Correct answer: (C) it decreases across a period. Ionization energy increases (not decreases) across a period from left to right.

**Which of the following statements about electron affinity is incorrect? (A) it is measured in kJ/mol (B) it involves release of energy (C) it decreases across a period (D) it decreases down a group**

Correct answer: (C) it decreases across a period. Electron affinity increases (not decreases) across a period from left to right.

## Quick Revision Summary

- Periodic law: properties of elements are periodic functions of their atomic numbers (Moseley, 1913)
- Modern periodic table: 7 periods (horizontal rows), 18 groups (vertical columns)
- Period 1 = short (2 elements); periods 2-3 = normal (8 each); periods 4-5 = long (18 each); periods 6-7 = very long (32 each)
- 4 blocks: s-block (groups 1-2), p-block (groups 13-18), d-block (groups 3-12, transition), f-block (lanthanides/actinides)
- Across a period (left→right): atomic size ↓, shielding ↓, ionization energy ↑, electron affinity ↑, electronegativity ↑
- Down a group (top→bottom): atomic size ↑, shielding ↑, ionization energy ↓, electron affinity ↓, electronegativity ↓
- Shielding effect = inner electrons reduce the nuclear pull felt by valence electrons; increases down a group

## Exam Tips

- Memorise the periodic law in both forms (Mendeleev's atomic-mass version and the modern atomic-number version) and be ready to explain why the modern version replaced it
- Learn the period names (short, normal, long, very long) with their exact element counts -- a common short-question topic

- Practice identifying which block (s, p, d, f) a given group number belongs to
- Build a single mental table of all 5 periodic trends (atomic size, shielding, ionization energy, electron affinity, electronegativity) and their direction across periods vs down groups -- this connects many exam questions
- Be ready to explain WHY a trend occurs, not just state it -- effective nuclear charge and shielding effect are the two reasons behind every trend in this unit
- Practice numerical examples using real values (e.g. period 2 atomic radii, ionization energies) since board exams often ask you to explain a specific given data trend

# Unit 4: Structure of Molecules

Atoms combine with each other to form molecules because every atom has a natural tendency to become more stable by attaining the electronic configuration of the nearest noble gas -- 2 electrons in the valence shell (duplet rule) or 8 electrons (octet rule). The force of attraction that holds atoms together once they combine is called a chemical bond.

This unit covers the four types of chemical bonds -- ionic, covalent (single, double, triple, and dative/coordinate), and metallic -- along with intermolecular forces such as dipole-dipole interactions and hydrogen bonding, and finally how the nature of bonding in a compound determines its physical properties (melting point, conductivity, solubility, and more).

## Learning Objectives

- Find the number of valence electrons in an atom using the Periodic Table
- State the octet and duplet rule and explain how elements attain stability
- Describe the formation of cations from metal atoms and anions from non-metal atoms
- Describe the characteristics of an ionic bond and identify properties of ionic compounds
- Describe the formation of a covalent bond, and distinguish single, double and triple covalent bonds with examples
- Explain dative (coordinate) covalent bonding, and polar vs non-polar covalent bonds
- Describe the formation of a metallic bond and explain the properties of metals
- Describe intermolecular forces (dipole-dipole, hydrogen bonding) and relate bonding type to compound properties

## Key Concepts

### 4.1 Why Do Atoms Form Chemical Bonds?

Atoms achieve stability by attaining the electronic configuration of the nearest noble gas -- 2 electrons in the valence shell (the duplet rule, for elements like H and He with only an s-subshell) or 8 electrons (the octet rule). Noble gases already have completely filled valence shells, so they neither gain, lose, nor share electrons, making them non-reactive; every other atom bonds with others in an attempt to copy this stable arrangement.

An atom can complete its valence shell in three ways: by giving away valence electrons (if it has fewer than 3), by gaining electrons from another atom (if it has 5 or more), or by sharing valence electrons with another atom. An atom's group number in the Periodic Table (based on its valence electron count) predicts which of these three routes it will take.

### 4.2 The Chemical Bond

A chemical bond is the force of attraction between atoms that holds them together in a substance. When two atoms approach each other, both attractive and repulsive forces act between them; if the net attractive forces dominate, the system's energy is lowered and a stable bond (and molecule) forms. If repulsive forces dominate instead, no bond forms. Bond formation between ions is due to electrostatic attraction; between similar (or comparably electronegative) atoms, it happens through the sharing of electrons.

### 4.3.1 Ionic Bond

Group 1 and 2 metals tend to lose their valence electrons to form positively charged cations, while Group 15-17 non-metals (electronegative, high electron affinity) tend to gain electrons to form negatively charged anions. An ionic bond forms when there is a complete transfer of one or more electrons from a metal atom to a non-metal atom, and the resulting oppositely charged ions are held together by electrostatic attraction.

In sodium chloride formation, sodium (1s2 2s2 2p6 3s1) loses its single valence electron to become Na+ (with a stable 8-electron outer shell), and chlorine (1s2

2s2 2p6 3s2 3p5) gains that electron to become $Cl^-$ (also 8 electrons outermost): $Na^+ + Cl^- \rightarrow NaCl$. Only valence electrons take part in ionic bonding, and heat is usually released during the reaction. Compounds formed this way are called ionic compounds.

### 4.3.2 Covalent Bond and Its Types

Elements of Groups 13-17 typically form bonds with each other by mutually sharing valence electrons rather than transferring them -- this is a covalent bond. The shared electrons are called bond pair electrons, and covalent bonds are classified by how many bond pairs are shared: a single covalent bond (one shared pair, e.g. $H_2$, $Cl_2$), a double covalent bond (two shared pairs, e.g. $O_2$, $C_2H_4$/ethene), and a triple covalent bond (three shared pairs, e.g. $N_2$, $C_2H_2$/ethyne). By sharing electrons this way, each atom attains an octet (or duplet) configuration. Lewis structures use dots or crosses around an element's symbol to represent valence electrons and bonding.

### 4.3.3 Dative (Coordinate) Covalent Bond

In a dative or coordinate covalent bond, the shared electron pair is contributed entirely by one atom (the donor), while the other atom (the acceptor) contributes none; a small arrow pointing from donor to acceptor represents this bond. A lone pair (a non-bonded electron pair, such as the one on nitrogen in ammonia, $NH_3$) can be donated -- for example, when $H^+$ approaches $NH_3$, nitrogen donates its lone pair to form the ammonium ion, $NH_4^+$. Boron trifluoride ($BF_3$) is electron-deficient (short of 2 electrons in boron's outer shell even after normal covalent bonding), so it readily accepts a lone pair from a donor molecule like ammonia to form a coordinate covalent bond.

### 4.3.4 Polar and Non-Polar Covalent Bonds

When a covalent bond forms between two identical atoms (e.g. $H_2$, $Cl_2$), the shared electron pair is attracted equally by both, producing a non-polar (pure) covalent bond. When it forms between two different atoms with different electronegativities, the more electronegative atom pulls the shared pair more strongly, creating a partial negative charge (delta-minus) on itself and a partial positive charge (delta-plus) on the other atom -- this is a polar covalent bond (e.g. water, HF, HCl). As a rule of thumb, if the electronegativity difference between two bonding atoms exceeds 1.7, the bond is predominantly ionic; if it is less than 1.7, the bond is predominantly covalent.

### 4.3.5 Metallic Bond

A metallic bond forms between metal atoms through mobile ('free') electrons. Because metal atoms are relatively large with more shells between the nucleus and valence electrons, and because they have low ionization energies, their outermost electrons detach easily and move freely throughout the metal rather than staying attached to any one atom. This creates a 'sea' of delocalized electrons surrounding a lattice of positive metal ions (nuclei), and it is this mobile electron sea that holds the metal atoms together and explains metals' characteristic properties -- electrical/thermal conductivity, malleability, ductility, and metallic luster.

## 4.4 Intermolecular Forces: Dipole-Dipole and Hydrogen Bonding

In addition to the strong chemical bonds within a molecule, weaker intermolecular forces (collectively called van der Waals forces) act between molecules. For example, breaking the intermolecular forces between liquid HCl molecules takes only about 17 kJ/mol, while breaking the H-Cl chemical bond itself takes about 430 kJ/mol -- intermolecular forces are far weaker than chemical bonds. Dipole-dipole interaction occurs when the partial positive end of one polar molecule (like HCl) is attracted to the partial negative end of a neighbouring molecule.

Hydrogen bonding is a special, stronger dipole-dipole attraction that occurs when a hydrogen atom is covalently bonded to a small, highly electronegative atom with lone pairs (N, O, or F) in one molecule, and is then attracted to a similar electronegative atom in a neighbouring molecule. Hydrogen bonding significantly raises boiling points (water's 100°C boiling point vs alcohol's 78°C is due to stronger/more hydrogen bonding in water) and explains why ice floats: as water freezes, hydrogen bonds arrange the molecules into an open, expanded structure, making ice (0.917 g/cm3) less dense than liquid water (1.00 g/cm3).

### 4.5 Nature of Bonding and Properties of Compounds

Ionic compounds consist of ions (not molecules) arranged in an orderly crystal lattice held by strong electrostatic forces. They are mostly crystalline solids with high melting/boiling points (e.g. NaCl melts at 800°C), show negligible electrical conductivity as solids but conduct well when molten or dissolved (due to free-moving ions), and dissolve readily in polar solvents like water.

Covalent compounds consist of molecules formed by shared electrons; the covalent bond is generally weaker than the ionic bond. They usually have low melting/boiling points, are usually poor conductors of electricity (except when a polar covalent compound dissolves in a polar solvent), are usually insoluble in water but soluble in non-aqueous solvents (benzene, ether, alcohol), and large 3D-bonded covalent crystals can be very hard with very high melting points. Non-polar covalent compounds generally do not dissolve in water or conduct electricity, while polar covalent compounds usually dissolve in water and their aqueous solutions can conduct electricity.

Metals show metallic luster, are malleable (can be rolled into sheets) and ductile (can be drawn into wires), usually have high melting/boiling points, form cations easily due to low ionization energy, and are good conductors of heat and electricity in both solid and liquid states because of their mobile electrons.

## Important Definitions

### What is the octet rule?

The tendency of an atom to attain 8 electrons in its outermost (valence) shell, either by losing, gaining, or sharing electrons, to achieve the stable electronic configuration of the nearest noble gas.

### What is the duplet rule?

The tendency of an atom (such as hydrogen or helium, which has only an s-subshell) to attain 2 electrons in its valence shell for stability.

### Define a chemical bond.

A force of attraction between atoms that holds them together in a substance.

### Define an ionic bond.

A chemical bond formed by the complete transfer of one or more electrons from one atom (usually a metal) to another (usually a non-metal), producing oppositely charged ions held together by electrostatic attraction.

### Define a covalent bond.

A chemical bond formed by the mutual sharing of valence electrons between two atoms, usually non-metals.

### What is a bond pair?

A pair of electrons shared between two atoms that forms a covalent bond.

### What is a lone pair?

A pair of valence electrons on an atom that is not involved in bonding, such as the lone pair on nitrogen in ammonia.

### Define a dative (coordinate) covalent bond.

A covalent bond in which the shared electron pair is contributed entirely by one atom (the donor) and accepted by the other atom (the acceptor).

### What is a polar covalent bond?

A covalent bond formed between two different atoms with different electronegativities, resulting in a partial positive charge on one atom and a partial negative charge on the other.

### Define a metallic bond.

A bond formed between metal atoms (positive ions) held together by a 'sea' of mobile, delocalized electrons.

### What are intermolecular forces?

Relatively weak forces of attraction that exist between molecules, in addition to the stronger chemical bonds within them.

### Define hydrogen bonding.

A strong type of dipole-dipole intermolecular attraction that occurs when a hydrogen atom bonded to a small, highly electronegative atom (N, O or F) is attracted to a similar electronegative atom in a neighbouring molecule.

### What is malleability?

The property of a metal that allows it to be rolled or hammered into thin sheets.

### What is ductility?

The property of a metal that allows it to be drawn into thin wires.

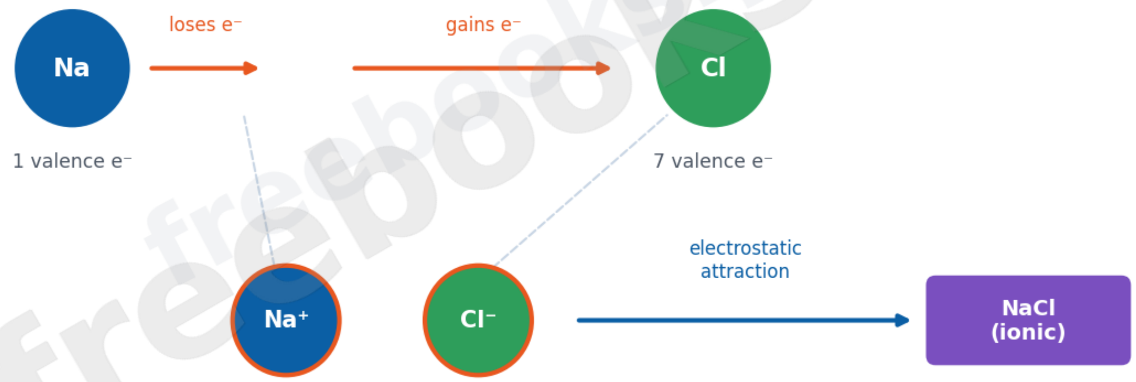
## Key Formulas

| Topic | Formula |
|---|---|
| Ionic bond formation (NaCl) | $Na^+ + Cl^- \rightarrow NaCl$ |
| Sodium electron loss | Na (1s2 2s2 2p6 3s1) → Na+ (1s2 2s2 2p6) + e- |
| Chlorine electron gain | Cl (1s2 2s2 2p6 3s2 3p5) + e- → Cl- (1s2 2s2 2p6 3s2 3p6) |
| Ionic vs covalent bond rule | Electronegativity difference > 1.7 → predominantly ionic; < 1.7 → predominantly covalent |
| Intermolecular force vs bond energy (HCl) | Intermolecular force ≈ 17 kJ/mol; H-Cl chemical bond ≈ 430 kJ/mol |
| Density comparison (water vs ice) | Liquid water at 0°C = 1.00 g/$cm^3$; Ice at 0°C = 0.917 g/$cm^3$ |

## Diagrams

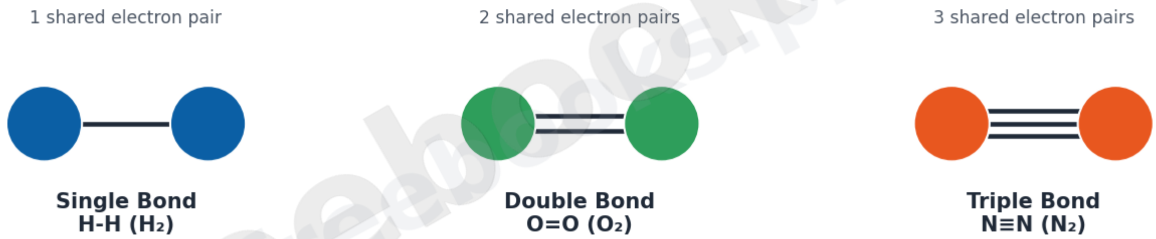
**Ionic Bond Formation: Na + Cl → NaCl:** Electron transfer from sodium to chlorine, producing Na+ and Cl- ions held together by electrostatic attraction.

**Ionic Bond Formation: Sodium + Chlorine → Sodium Chloride**



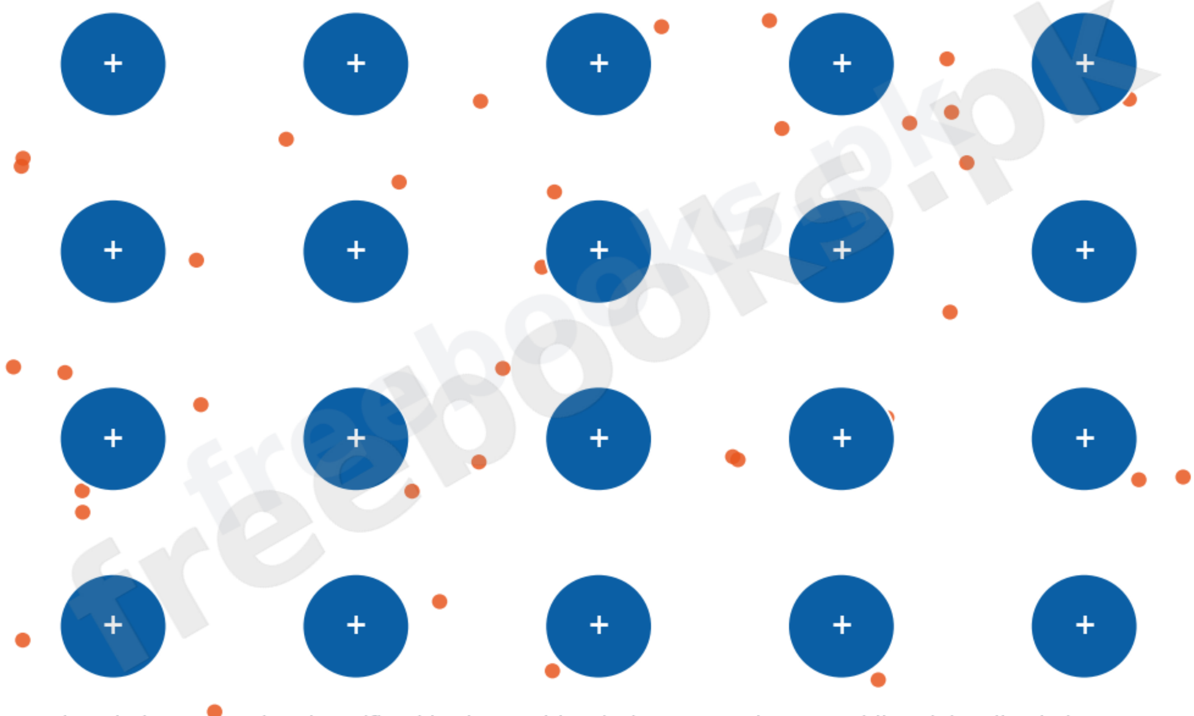

**Types of Covalent Bonds:** Single (H2), double (O2) and triple (N2) covalent bonds shown by the number of shared electron pairs.

**Types of Covalent Bonds by Number of Shared Electron Pairs**




**Metallic Bonding: Sea of Free Electrons:** Positive metal ions arranged in a lattice, surrounded by mobile, delocalized electrons that hold the metal together.

**Metallic Bonding: Positive Ions in a 'Sea' of Free Electrons**



*Blue circles = metal cations (fixed lattice positions) | Orange dots = mobile / delocalized electrons*

## Short Questions & Answers

### Why do atoms react with other atoms?

Atoms react to attain a stable electronic configuration, similar to that of the nearest noble gas, by losing, gaining, or sharing valence electrons.

### Why does sodium form Na+ and not Na2+ or Na-?

Sodium has only 1 valence electron; losing that single electron gives it the stable, completely filled configuration of neon (8 electrons in its new outermost shell), so it forms Na+.

### What distinguishes a single, double, and triple covalent bond?

A single bond has 1 shared electron pair (e.g. H2), a double bond has 2 shared pairs (e.g. O2), and a triple bond has 3 shared pairs (e.g. N2).

### Why is BF3 described as electron-deficient?

Even after forming three covalent bonds with fluorine, boron's outermost shell has only 6 electrons, 2 short of a full octet, making it eager to accept a lone pair from a donor atom.

### Why is HCl a polar molecule?

Because chlorine is more electronegative than hydrogen, it attracts the shared electron pair more strongly, creating a partial negative charge on chlorine and a partial positive charge on hydrogen.

### Why do metals conduct electricity well?

Because metal atoms release their outermost electrons into a mobile 'sea' of delocalized electrons that can move freely throughout the metal, carrying electric charge.

### Why are intermolecular forces much weaker than chemical bonds?

Intermolecular forces arise from temporary or partial charge attractions between molecules, while chemical bonds involve actual transfer or sharing of electrons between atoms -- for HCl, breaking the intermolecular force takes ~17 kJ/mol versus ~430 kJ/mol to break the chemical bond.

### Why does ice float on water?

Hydrogen bonds arrange water molecules into an open, expanded structure as they freeze, making solid ice less dense (0.917 g/cm3) than liquid water (1.00 g/cm3).

### Why do ionic compounds conduct electricity only when molten or dissolved?

In the solid state, ions are fixed in a rigid crystal lattice and cannot move, but when molten or dissolved, the ions become free to move and carry electric current.

### Why are covalent compounds usually poor conductors of electricity?

Because they consist of neutral molecules rather than free-moving charged ions or electrons, so they generally have no charge carriers available to conduct current.

## Long Questions & Answers

### Define an ionic bond and describe, with a worked example, how it forms between two atoms.

#### What is an ionic bond?

An ionic bond is a chemical bond formed by the complete transfer of one or more electrons from one atom (typically a metal with few valence electrons) to another atom (typically a non-metal with many valence electrons), producing oppositely charged ions that are held together by electrostatic attraction.

#### How does sodium change during ionic bond formation?

Sodium (11Na, configuration 1s2 2s2 2p6 3s1) has just 1 electron in its valence shell. By losing this single electron, sodium attains the stable configuration of neon (8 electrons in its new outermost shell) and becomes a positively charged Na+ ion.

#### How does chlorine change during ionic bond formation?

Chlorine (17Cl, configuration 1s2 2s2 2p6 3s2 3p5) has 7 electrons in its valence shell. By gaining 1 electron from sodium, chlorine completes its octet (8 electrons) and becomes a negatively charged Cl- ion.

#### How do the resulting ions form the final compound?

Once formed, the oppositely charged Na+ and Cl- ions attract each other strongly through electrostatic force, combining as Na+ + Cl- → NaCl. Only valence-shell electrons take part in this process, and heat is usually released, forming a stable ionic compound.

**Explain the three main types of covalent bonds and the dative covalent bond, with one example of each.**

**What is a single covalent bond?**

A single covalent bond forms when each bonded atom contributes one electron, creating one shared (bond) pair, shown as a single line between atoms. Hydrogen gas ($H_2$, H-H) is a classic example.

**What is a double covalent bond?**

A double covalent bond forms when each bonded atom contributes two electrons, creating two shared pairs, shown as a double line between atoms. Oxygen gas ($O_2$, O=O) is an example.

**What is a triple covalent bond?**

A triple covalent bond forms when each bonded atom contributes three electrons, creating three shared pairs, shown as a triple line between atoms. Nitrogen gas ($N_2$, N≡N) is an example.

**What is a dative (coordinate) covalent bond, and how does it differ from the others?**

In a dative covalent bond, only one atom (the donor) contributes both electrons of the shared pair, while the other atom (the acceptor) contributes none -- unlike ordinary covalent bonds where both atoms contribute one electron each. For example, nitrogen's lone pair in ammonia ($NH_3$) is donated to $H^+$ to form the ammonium ion ($NH_4^+$), and ammonia can similarly donate its lone pair to electron-deficient $BF_3$.

## Multiple Choice Questions (MCQs)

### Atoms react with each other because: (A) they are attracted to each other (B) they are short of electrons (C) they want to attain stability (D) they want to disperse

Correct answer: (C) they want to attain stability. Atoms react in order to attain a stable, noble-gas-like electronic configuration.

**An atom having six electrons in its valence shell will achieve noble gas configuration by: (A) gaining one electron (B) losing all electrons (C) gaining two electrons (D) losing two electrons**

Correct answer: (C) gaining two electrons. An atom with 6 valence electrons needs 2 more to complete its octet, so it gains 2 electrons.

**The octet rule refers to: (A) a description of eight electrons (B) a picture of electronic configuration (C) a pattern of electronic configuration (D) the attaining of eight electrons in the valence shell**

Correct answer: (D) the attaining of eight electrons in the valence shell. The octet rule refers to an atom attaining 8 electrons in its outermost (valence) shell for stability.

**Transfer of electrons between atoms results in: (A) metallic bonding (B) ionic bonding (C) covalent bonding (D) coordinate covalent bonding**

Correct answer: (B) ionic bonding. Complete transfer of electrons between atoms produces an ionic bond.

**When an electronegative element combines with an electropositive element, the bonding type is: (A) covalent (B) ionic (C) polar covalent (D) coordinate covalent**

Correct answer: (B) ionic. A large electronegativity difference between an electropositive metal and an electronegative non-metal produces an ionic bond.

**A bond formed between two non-metals is expected to be: (A) covalent (B) ionic (C) coordinate covalent (D) metallic**

Correct answer: (A) covalent. Two non-metals with comparable electronegativities typically form a covalent bond by sharing electrons.

**A bond pair in covalent molecules usually consists of: (A) one electron (B) two electrons (C) three electrons (D) four electrons**

Correct answer: (B) two electrons. A bond pair is a shared pair of two electrons between two bonded atoms.

**Ice floats on water because: (A) ice is denser than water (B) ice is crystalline in nature (C) water is denser than ice (D) water molecules move randomly**

Correct answer: (C) water is denser than ice. Liquid water (1.00 g/cm3) is denser than ice (0.917 g/cm3) due to the open, expanded hydrogen-bonded structure of ice.

**A triple covalent bond involves how many shared electrons in total? (A) eight (B) six (C) four (D) only three**

Correct answer: (B) six. A triple covalent bond involves 3 shared pairs of electrons, totalling 6 electrons.

**Which one of the following is the weakest force among the choices given? (A) ionic force (B) metallic force (C) intermolecular force (D) covalent force**

Correct answer: (C) intermolecular force. Intermolecular forces are much weaker than ionic, metallic, or covalent chemical bonds.

## Quick Revision Summary

- Atoms bond to attain noble gas configuration: duplet rule (2 e-, H/He) or octet rule (8 e-, most elements)
- Ionic bond = complete electron transfer (metal → non-metal), held by electrostatic attraction, e.g. NaCl
- Covalent bond = mutual sharing; single (1 pair), double (2 pairs), triple (3 pairs) bonds
- Dative/coordinate covalent bond = both shared electrons come from ONE atom (the donor), e.g. $NH_3 \rightarrow NH_4^+$
- Polar covalent bond forms when electronegativity difference exists; ionic if difference > 1.7, covalent if < 1.7
- Metallic bond = positive metal ions held together by a 'sea' of mobile, delocalized electrons
- Intermolecular forces (dipole-dipole, hydrogen bonding) are much weaker than chemical bonds
- Hydrogen bonding explains water's high boiling point and why ice floats (less dense than liquid water)
- Ionic compounds: high melting/boiling points, conduct only when molten/dissolved, soluble in water
- Covalent compounds: low melting/boiling points, usually poor conductors, often insoluble in water

## Exam Tips

- Practice drawing Lewis (dot-and-cross) structures for common molecules: $H_2$, $O_2$, $N_2$, $NH_3$, $CH_4$, $H_2O$
- Memorise which groups tend to lose electrons (1, 2) and which tend to gain electrons (15-17)
- Learn the 1.7 electronegativity-difference rule for predicting ionic vs covalent bonding -- a common numerical-reasoning question
- Be ready to explain the exact process of ionic bond formation for a given pair of elements, step by step (electron loss, electron gain, then attraction)
- Understand the donor/acceptor concept for dative bonds -- a frequently confused topic; remember only ONE atom contributes both electrons
- Connect bonding type to properties: if asked why a compound has a certain property (high melting point, poor conductivity, etc.), trace it back to the type of bond present

# Unit 5: Physical States of Matter

Matter exists in three physical states -- gas, liquid, and solid -- distinguished mainly by the strength of intermolecular forces between their particles. Gases have very weak intermolecular forces and no definite shape or volume; liquids have moderate forces, giving them a definite volume but no fixed shape; solids have strong forces, giving them both definite shape and volume.

This unit covers the typical properties of gases (diffusion, effusion, pressure, compressibility), the two gas laws (Boyle's Law and Charles's Law), the typical properties of liquids (evaporation, vapour pressure, boiling point, freezing point, diffusion, density), and the typical properties of solids (melting point, rigidity, density, amorphous vs crystalline solids, and allotropy).

## Learning Objectives

- Describe the effect of pressure and temperature changes on the volume of a gas
- Compare the three physical states of matter with regard to intermolecular forces
- Account for pressure-volume changes in a gas using Boyle's Law
- Account for temperature-volume changes in a gas using Charles's Law
- Explain the properties of gases: diffusion, effusion, and pressure
- Explain the properties of liquids: evaporation, vapour pressure, and boiling point
- Explain the effect of temperature and external pressure on vapour pressure and boiling point
- Describe the physical properties of solids (melting point) and differentiate amorphous from crystalline solids
- Explain the allotropic forms of elements

## Key Concepts

### 5.1 Typical Properties of Gases

Diffusion is the spontaneous mixing of gas molecules by random motion and collisions to form a homogeneous mixture; its rate depends on molecular mass, with lighter gases diffusing faster ($H_2$ diffuses four times faster than $O_2$). Effusion is the escape of gas molecules through a tiny hole into a region of lower pressure, and like diffusion, lighter gases effuse faster than heavier ones.

Pressure (P = Force/Area) is exerted by gas molecules constantly colliding with container walls; its SI unit is the Pascal (Pa = 1 $N/m^2$). Standard atmospheric pressure is the pressure exerted by a 760 mm column of mercury at sea level: 1 atm = 760 mmHg = 760 torr = 101325 Pa. Gases are highly compressible (due to empty space between molecules), highly mobile (due to high kinetic energy), and have low density compared to liquids and solids -- gas density is measured in $g/dm^3$ (liquids/solids are 1000 times denser and measured in $g/cm^3$), and gas density increases on cooling as volume decreases.

### 5.2.1 Boyle's Law

Boyle's Law (1662) states that the volume of a given mass of gas is inversely proportional to its pressure at constant temperature: $V \propto 1/P$, or $PV = k$ (a constant). For two states of the same gas sample, $P_1V_1 = P_2V_2$. This was verified experimentally by compressing a gas in a cylinder: at 2 atm the volume was 1 $dm^3$, at 4 atm it was 0.5 $dm^3$, at 6 atm it was 0.33 $dm^3$, and at 8 atm it was 0.25 $dm^3$ -- in every case, the product $P \times V$ remained constant at 2 atm·$dm^3$.

### 5.2.2 Absolute Temperature Scale and Charles's Law

The Kelvin (absolute) scale starts at 0 K (-273.15°C), called absolute zero -- the temperature at which an ideal gas would theoretically have zero volume. Conversion: K = °C + 273, and °C = K - 273. Charles's Law (1787) states that the volume of a given mass of gas is directly proportional to its absolute temperature at constant pressure: $V \propto T$, or $V/T = k$. For two states, $V_1/T_1 = V_2/T_2$. This was verified experimentally: heating a gas from 25°C ($V_1$ = 50 $cm^3$) to 100°C increased its volume to about 62.5 $cm^3$.

## Physical States and Intermolecular Forces

In the gaseous state, molecules are far apart, so intermolecular forces are very weak. In the liquid state, molecules are much closer together, developing stronger intermolecular forces that affect diffusion, evaporation, vapour pressure, and boiling point -- compounds with stronger intermolecular forces have higher boiling points. In the solid state, intermolecular forces dominate so strongly that molecules appear almost motionless, arranging in a regular pattern that makes solids denser than liquids or gases.

## 5.3 Typical Properties of Liquids

Evaporation is the endothermic process of a liquid changing into vapour (e.g. converting 1 mole of liquid water to vapour requires 40.7 kJ). It occurs because a fraction of molecules have more than average kinetic energy and escape the liquid surface; it is a cooling process (remaining molecules lose energy) and depends on surface area (more surface area, more evaporation), temperature (higher temperature, faster evaporation), and intermolecular forces (stronger forces, slower evaporation -- e.g. alcohol evaporates faster than water).

Vapour pressure is the pressure exerted by a liquid's vapour when it is in dynamic equilibrium with the liquid (rate of evaporation = rate of condensation) at a given temperature. It depends on the nature of the liquid (polar liquids have lower vapour pressure than non-polar ones due to stronger intermolecular forces), the size of molecules (smaller molecules evaporate more easily, giving higher vapour pressure), and temperature (higher temperature increases vapour pressure).

Boiling point is the temperature at which a liquid's vapour pressure becomes equal to atmospheric (or external) pressure. It depends on the nature of the liquid (polar liquids have higher boiling points), intermolecular forces (stronger forces mean higher boiling points), and external pressure (higher external pressure raises the boiling point, as used in a pressure cooker). Freezing point is the temperature at which the vapour pressures of the liquid and solid phases become equal, and the two phases coexist in dynamic equilibrium.

Liquids also diffuse (much more slowly than gases), with diffusion depending on intermolecular forces (weaker forces diffuse faster), molecule size (bigger molecules diffuse slower, e.g. honey diffuses slower than alcohol in water), molecule shape (regular shapes diffuse faster), and temperature (higher temperature speeds up diffusion). Liquids are denser than gases because their molecules are closely packed with negligible space between them (e.g. water is 1.0 g/cm3 versus air at 0.001 g/cm3).

### 5.4 Typical Properties of Solids

The melting point is the temperature at which a solid, upon heating, coexists in dynamic equilibrium with its liquid state; ionic and covalent network solids (macromolecules) have very high melting points. Solids are rigid because their particles occupy fixed positions and can only vibrate, not move freely. Solids are the densest state of matter because their particles are closely packed with no empty space between them (e.g. aluminium 2.70 g/cm3, iron 7.86 g/cm3, gold 19.3 g/cm3).

### 5.5 Types of Solids: Amorphous and Crystalline

Amorphous solids (meaning 'shapeless') have particles that are not regularly arranged, so they lack a sharp melting point -- examples include plastic, rubber, and glass. Crystalline solids have particles arranged in a definite, repeating three-dimensional pattern with well-defined faces meeting at fixed angles, giving them sharp melting points -- examples include diamond and sodium chloride.

### 5.6 Allotropy

Allotropy is the existence of an element in more than one form within the same physical state, arising either from different numbers of atoms per molecule (e.g. oxygen O2 and ozone O3) or from different arrangements of atoms/molecules within a crystal (e.g. sulphur's S8 molecules arranged differently in rhombic vs monoclinic crystals). Allotropes of the same element always show different physical properties but identical chemical properties. The temperature at which one allotrope converts to another is called the transition temperature -- for example, rhombic sulphur converts to monoclinic sulphur at 96°C, white phosphorus converts to red phosphorus at 250°C, and grey (cubic) tin converts to white (tetragonal) tin at 13.2°C.

## Important Definitions

### Define diffusion.

The spontaneous mixing up of molecules by random motion and collisions to form a homogeneous mixture.

### Define effusion.

The escape of gas molecules through a tiny hole into a space of lesser pressure.

### Define pressure and give its SI unit.

Pressure is force exerted per unit surface area (P = F/A); its SI unit is the Pascal (Pa), equal to 1 N/m2.

### Define standard atmospheric pressure.

The pressure exerted by a mercury column of 760 mm height at sea level, equal to 1 atm = 760 mmHg = 101325 Pa.

### State Boyle's Law.

The volume of a given mass of a gas is inversely proportional to its pressure at constant temperature (PV = constant).

### State Charles's Law.

The volume of a given mass of a gas is directly proportional to its absolute temperature at constant pressure (V/T = constant).

### What is absolute zero?

0 K or -273.15°C, the temperature at which an ideal gas would theoretically have zero volume.

### Define evaporation.

The process of a liquid changing into a gas (vapour) phase; it is an endothermic (cooling) process.

### Define vapour pressure.

The pressure exerted by the vapours of a liquid when they are in dynamic equilibrium with the liquid at a particular temperature.

### Define boiling point.

The temperature at which the vapour pressure of a liquid becomes equal to the atmospheric (or external) pressure.

### Define freezing point.

The temperature at which the vapour pressures of the liquid and solid states of a substance become equal, and the two phases coexist in dynamic equilibrium.

### Define melting point.

The temperature at which a solid, upon heating, starts to melt and coexists in dynamic equilibrium with its liquid state.

### Differentiate amorphous and crystalline solids in one line each.

Amorphous solids have irregularly arranged particles and no sharp melting point (e.g. glass); crystalline solids have a definite 3D particle arrangement and a sharp melting point (e.g. diamond).

### Define allotropy.

The existence of an element in more than one form within the same physical state, showing different physical but identical chemical properties.

### Define transition temperature.

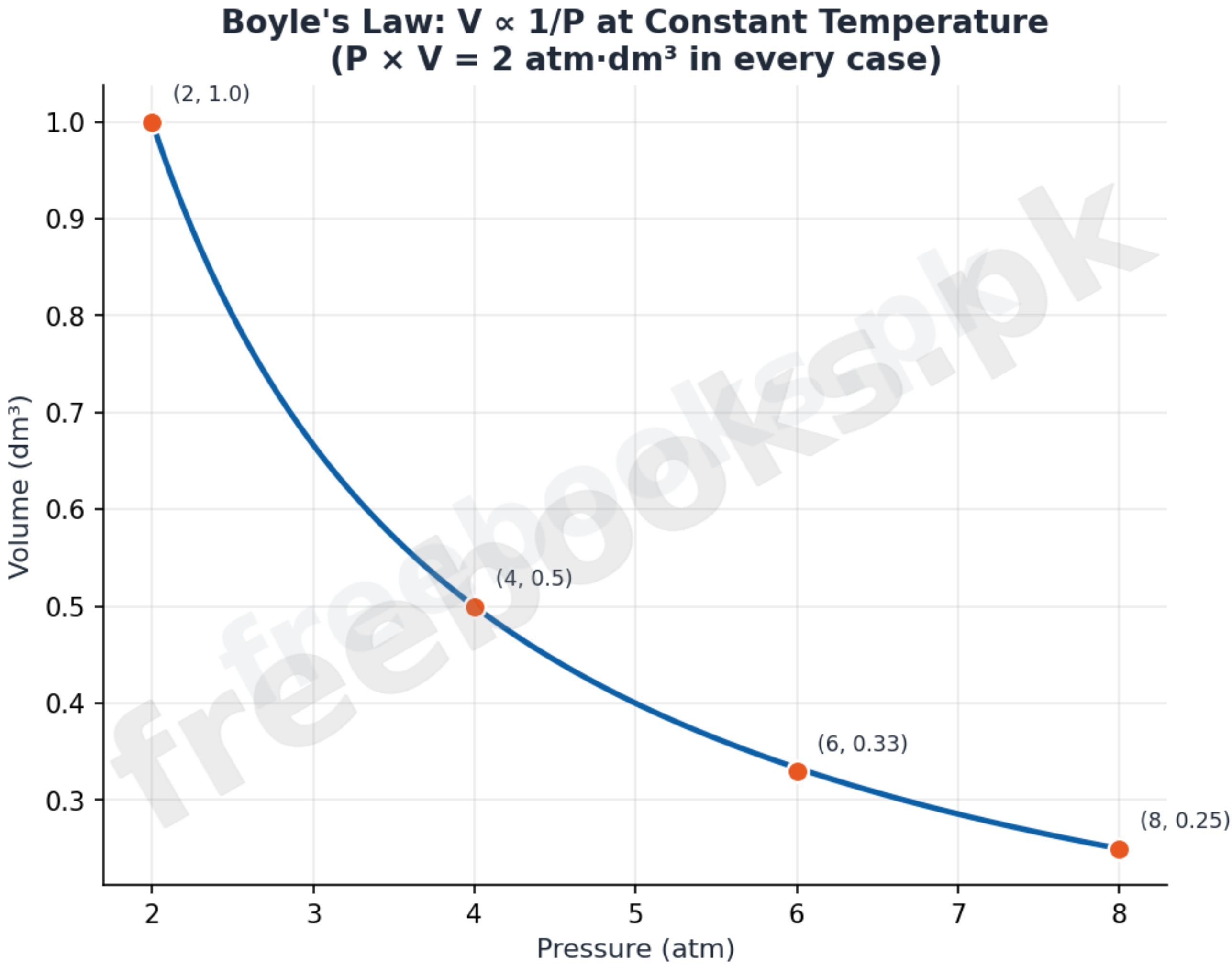
The temperature at which one allotropic form of an element converts into another.

## Key Formulas

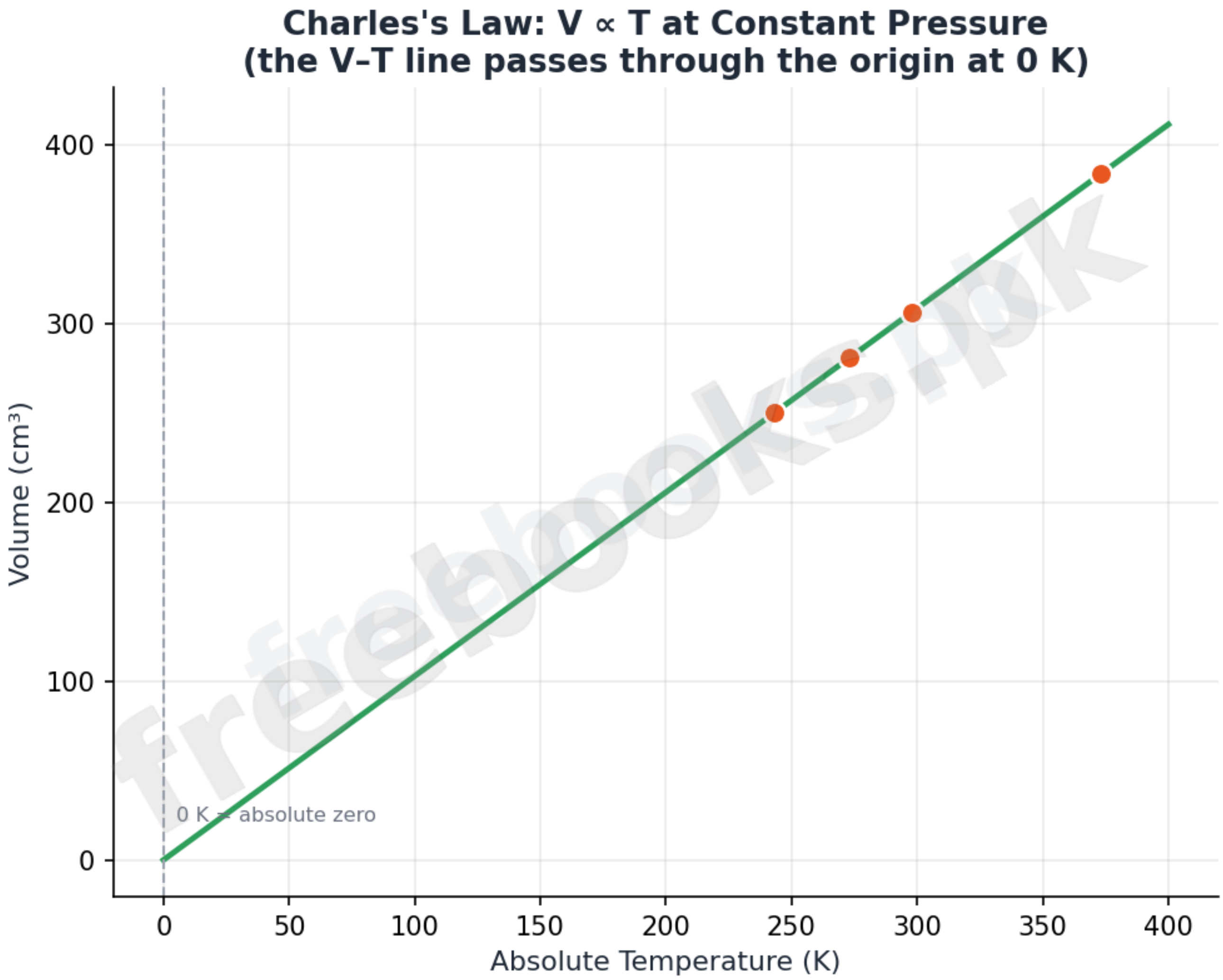
| Topic | Formula |
|---|---|
| Pressure | P = Force (F) / Area (A); 1 Pascal (Pa) = 1 N $m^{-2}$ |
| Standard atmospheric pressure | 1 atm = 760 mmHg = 760 torr = 101325 Pa |
| Boyle's Law | P1 V1 = P2 V2 (constant temperature) |
| Kelvin-Celsius conversion | K = °C + 273; °C = K - 273 |
| Charles's Law | V1 / T1 = V2 / T2 (constant pressure) |
| Evaporation of water (per mole) | H2O(l) → H2O(g), ΔH°vap = +40.7 kJ $mol^{-1}$ |

## Diagrams

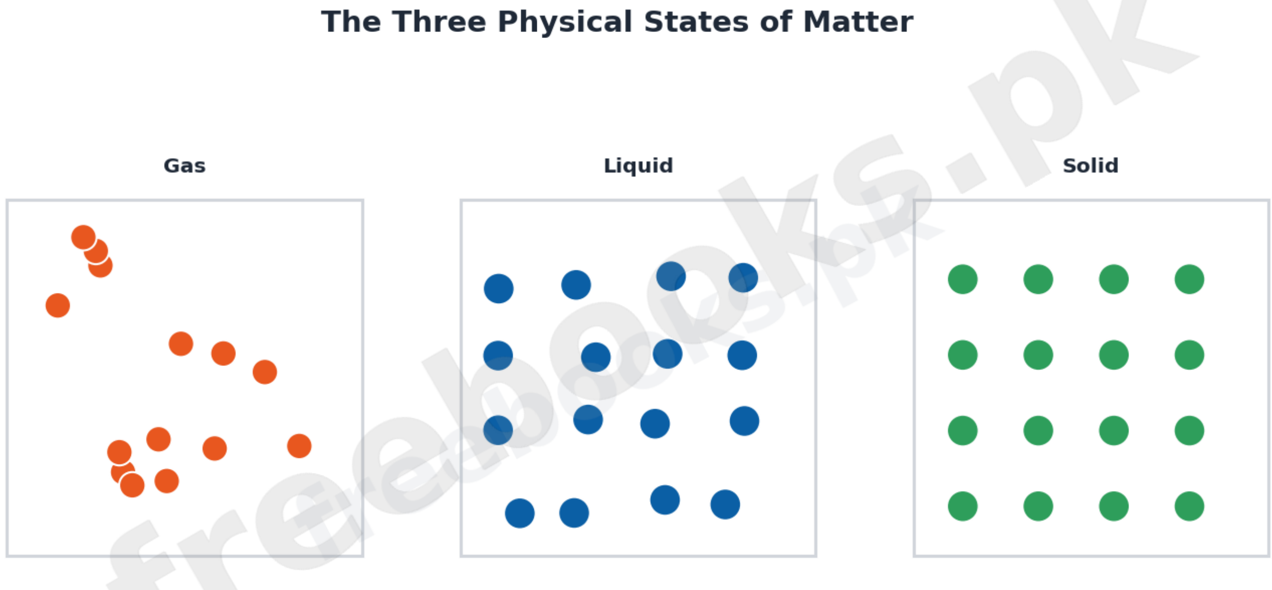
**Boyle's Law: V vs P:** Graph of the experimental Boyle's Law data (2 atm/1 $dm^3$, 4 atm/0.5 $dm^3$, 6 atm/0.33 $dm^3$, 8 atm/0.25 $dm^3$) showing the inverse relationship between pressure and volume.




**Charles's Law: V vs T:** Graph showing volume increasing linearly with absolute temperature at constant pressure, extrapolating to zero volume at 0 K (absolute zero).




**The Three Physical States of Matter:** Molecular arrangement comparison of gas (far apart, random), liquid (close, can flow), and solid (fixed, regular pattern) showing increasing intermolecular force strength.



## Short Questions & Answers

### Why does hydrogen gas diffuse faster than oxygen gas?

Because diffusion rate depends on molecular mass, and hydrogen (lighter) diffuses about four times faster than the heavier oxygen molecule.

### Why are gases highly compressible?

Because there is a large amount of empty space between gas molecules, allowing them to be pushed much closer together under pressure.

### Why does the density of a gas increase on cooling?

Because cooling reduces the gas's volume while its mass stays the same, so density (mass/volume) increases.

### According to Boyle's Law, what happens to volume if pressure is doubled at constant temperature?

The volume is halved, since volume is inversely proportional to pressure at constant temperature.

### Why is evaporation considered a cooling process?

Because the molecules with the highest kinetic energy escape the liquid, leaving behind molecules with lower average kinetic energy, which lowers the liquid's (and its surroundings') temperature.

### Why does alcohol evaporate faster than water?

Because alcohol has weaker intermolecular forces than water, so its molecules can more easily overcome those forces and escape into vapour.

### Why does increasing external pressure raise a liquid's boiling point?

Because the liquid's vapour pressure must rise to a higher value to match the increased external pressure before boiling can begin, requiring a higher temperature.

### Why are solids denser than liquids and gases?

Because solid particles are closely packed with no empty spaces between them, unlike the more spread-out particles in liquids and especially gases.

### Why do amorphous solids not have a sharp melting point?

Because their particles are not arranged in a regular repeating pattern, so different parts of the solid melt at slightly different temperatures rather than all at once.

**Why do oxygen ($O_2$) and ozone ($O_3$) count as allotropes?**
Because they are two different molecular forms of the same element, oxygen, existing in the same physical state (gas) but with a different number of atoms per molecule.

## Long Questions & Answers

### State Boyle's Law and Charles's Law, and explain how each was experimentally verified.

#### What does Boyle's Law state, and how is it expressed?

Boyle's Law states that the volume of a given mass of a gas is inversely proportional to its pressure at constant temperature: $V \propto 1/P$, or $PV = k$ (a constant). For two states of the same gas, this gives $P_1V_1 = P_2V_2$.

#### How was Boyle's Law experimentally verified?

A fixed mass of gas was compressed in a cylinder with a movable piston: at 2 atm the volume was 1 $dm^3$, at 4 atm it fell to 0.5 $dm^3$, at 6 atm to 0.33 $dm^3$, and at 8 atm to 0.25 $dm^3$. In every case, the product of pressure and volume remained constant at 2 atm·$dm^3$, confirming the law.

#### What does Charles's Law state, and how is it expressed?

Charles's Law states that the volume of a given mass of a gas is directly proportional to its absolute temperature at constant pressure: $V \propto T$, or $V/T = k$. For two states, this gives $V_1/T_1 = V_2/T_2$.

#### How was Charles's Law experimentally verified?

A fixed mass of gas enclosed in a cylinder with a movable piston had an initial volume of 50 $cm^3$ at 25°C; when heated to 100°C, its volume increased to about 62.5 $cm^3$, demonstrating that volume increases proportionally with absolute temperature at constant pressure.

**Explain vapour pressure and boiling point, and describe the factors that affect each.**

#### What is vapour pressure, and when does it apply?

Vapour pressure is the pressure exerted by a liquid's vapour when the liquid and its vapour are in dynamic equilibrium at a given temperature -- that is, when the rate of evaporation equals the rate of condensation.

#### What factors affect vapour pressure?

Vapour pressure depends on the nature of the liquid (polar liquids have lower vapour pressure than non-polar liquids due to stronger intermolecular forces), the size of the molecules (smaller molecules evaporate more easily, giving higher vapour pressure), and temperature (vapour pressure increases as temperature rises).

#### What is boiling point, and how does it relate to vapour pressure?

Boiling point is the temperature at which a liquid's vapour pressure becomes equal to the atmospheric (or external) pressure -- at this point, vapourization occurs throughout the liquid, not just at its surface.

#### What factors affect boiling point?

Boiling point depends on the nature of the liquid (polar liquids generally have higher boiling points), intermolecular forces (stronger forces require a higher temperature to reach the necessary vapour pressure, giving a higher boiling point), and external pressure (increasing external pressure raises the boiling point, which is the principle behind a pressure cooker).

## Multiple Choice Questions (MCQs)

### How many times are liquids denser than gases (approximately)? (A) 100 times (B) 1000 times (C) 10,000 times (D) 100,000 times

Correct answer: (B) 1000 times. Liquids and solids are approximately 1000 times denser than gases, which is why their densities are measured in different units (g/cm3 vs g/dm3).

**Gas densities are typically expressed in units of: (A) mg $cm^{-3}$ (B) g $cm^{-3}$ (C) g $dm^{-3}$ (D) kg $dm^{-3}$**

Correct answer: (C) g $dm^{-3}$. Because gases are much less dense than liquids or solids, their densities are conventionally expressed in grams per cubic decimetre (g $dm^{-3}$).

**At the freezing point, which of the following coexist in dynamic equilibrium? (A) gas and solid (B) liquid and gas (C) liquid and solid (D) all of these**

Correct answer: (C) liquid and solid. At the freezing point, the liquid and solid phases of a substance coexist in dynamic equilibrium.

**Solid particles possess which type of motion? (A) rotational motion only (B) vibrational motion (C) translational motion only (D) both translational and vibrational motion**

Correct answer: (B) vibrational motion. Solid particles are fixed in position and can only vibrate about their mean position; they cannot move translationally.

**Which one of the following is NOT an amorphous solid? (A) rubber (B) plastic (C) glass (D) glucose**

Correct answer: (D) glucose. Glucose is a crystalline solid with a definite, regular particle arrangement and a sharp melting point, unlike rubber, plastic, and glass.

**One standard atmospheric pressure is equal to how many Pascals? (A) 101325 (B) 10325 (C) 106075 (D) 10523**

Correct answer: (A) 101325. 1 atm = 101325 Pa, by definition of the standard atmosphere.

**Which one of the following gases diffuses fastest? (A) hydrogen (B) helium (C) fluorine (D) chlorine**

Correct answer: (A) hydrogen. Hydrogen has the lowest molecular mass of the choices, so it diffuses fastest according to the relationship between diffusion rate and molecular mass.

**Which one of the following does NOT affect a liquid's boiling point? (A) intermolecular forces (B) external pressure (C) nature of the liquid (D) initial temperature of the liquid**

Correct answer: (D) initial temperature of the liquid. Boiling point depends on intermolecular forces, external pressure, and the nature of the liquid, but not on what temperature the liquid started at before heating.

**The density of a gas increases when its: (A) temperature is increased (B) pressure is increased (volume decreases) (C) volume is increased (D) none of these**

Correct answer: (B) pressure is increased (volume decreases). Increasing pressure at constant temperature decreases a gas's volume, which increases its density (mass/volume).

**The vapour pressure of a liquid increases with: (A) increase of external pressure (B) increase of temperature (C) increase of intermolecular forces (D) increase of polarity of molecules**

Correct answer: (B) increase of temperature. Vapour pressure rises as temperature increases, because more molecules gain enough kinetic energy to escape into the vapour phase.

## Quick Revision Summary

- 3 states of matter differ mainly by intermolecular force strength: gas (weakest) < liquid < solid (strongest)
- Diffusion/effusion rate depends on molecular mass -- lighter gases move faster
- Pressure: P = F/A; 1 Pa = 1 N/m$^2$; 1 atm = 760 mmHg = 101325 Pa
- Boyle's Law: PV = constant (T fixed) -- P1V1 = P2V2
- Charles's Law: V/T = constant (P fixed) -- V1/T1 = V2/T2; always convert °C to K first (K = °C + 273)
- Evaporation is endothermic/cooling; depends on surface area, temperature, intermolecular forces
- Vapour pressure = pressure of vapour in dynamic equilibrium with its liquid
- Boiling point = temperature where vapour pressure equals external pressure
- Solids: amorphous (no sharp m.p., e.g. glass) vs crystalline (sharp m.p., e.g. diamond, NaCl)
- Allotropy = same element, different forms, same state (e.g. O2/O3, rhombic/monoclinic sulphur)

## Exam Tips

- Always convert temperature to Kelvin (K = °C + 273) before using Charles's Law in any calculation
- Practice unit conversions between atm, mmHg, torr, and Pa -- these appear frequently in numerical questions
- Memorise the exact wording of Boyle's Law and Charles's Law, including which variable is held constant in each
- Be ready to explain WHY evaporation causes cooling -- a common short-question topic
- Learn the three factors affecting vapour pressure/boiling point (nature of liquid, intermolecular forces, temperature/pressure) and be ready to apply them to a specific example
- Practice distinguishing amorphous vs crystalline solids with the correct examples (glass/rubber/plastic vs diamond/NaCl)
- Know at least one worked example of allotropy (sulphur, phosphorus, or tin) with its transition temperature

# Unit 6: Solutions

A solution is a homogeneous mixture of two or more substances whose components cannot be visually distinguished from each other -- it exists as a single phase. Solutions can be gaseous (air), liquid (sea water), or solid (alloys like brass), depending on the physical state of the solvent. Of the nine possible solute-solvent combinations, liquid solutions are the most common because water is the most widely available solvent.

This unit covers the definitions of solution, aqueous solution, solute, and solvent; the distinction between saturated, unsaturated, and supersaturated solutions and the meaning of dilution; the nine types of solutions by physical state; the various units used to express concentration (percentage forms and molarity), including worked numerical problems; the factors affecting solubility (the 'like dissolves like' principle, solute-solvent interaction, and temperature); and the comparison of true solutions, colloids, and suspensions.

## Learning Objectives

- Define solution, aqueous solution, solute, and solvent, and give an example of each
- Explain the difference between saturated, unsaturated, and supersaturated solutions
- Describe the nine types of solutions formed from gas, liquid, and solid solutes and solvents, with examples
- Explain the concept of the concentration of a solution
- Define molarity and percentage solution, and solve numerical problems involving molarity
- Describe how dilute solutions are prepared from concentrated solutions of known molarity
- Apply the principle 'like dissolves like' to predict the solubility of one substance in another
- Explain the effect of temperature on the solubility of different solutes

- Compare solutions, colloids, and suspensions based on particle size, stability, and the Tyndall effect

## Key Concepts

### 6.1 Solution, Aqueous Solution, Solute, and Solvent

A solution is a homogeneous mixture of two or more substances in which the boundaries of the components cannot be distinguished -- it exists as a single phase. Examples include air (a solution of gases), brass (a solid solution of zinc and copper), and sugar dissolved in water (a liquid solution). Evaporation distinguishes a solution from a pure liquid: a pure liquid evaporates completely leaving no residue, while a solution leaves a residue behind. Alloys like brass are considered mixtures, not compounds, because they show the properties of their components and have variable composition.

An aqueous solution is formed by dissolving a substance in water, where water is present in the greater amount and acts as the solvent -- for example, sugar in water or table salt in water. Water is called the universal solvent because it dissolves the majority of compounds found in the earth's crust. The solute is the component present in smaller quantity (e.g. salt in salt solution), while the solvent is the component present in larger quantity and always dissolves the solute(s) -- in soft drinks, water is the solvent while sugar, salts, and $CO_2$ are all solutes.

### 6.2 Saturated, Unsaturated, and Supersaturated Solutions

A saturated solution contains the maximum amount of solute that can dissolve at a given temperature; at the particle level, undissolved solute exists in dynamic equilibrium with dissolved solute (solute (crystallized) $\rightleftharpoons$ solute (dissolved)) -- dissolution and crystallization continue, but the net dissolved amount stays constant. An unsaturated solution contains less solute than is needed to saturate it at that temperature and can still dissolve more.

A supersaturated solution is more concentrated than a saturated solution and is prepared by dissolving solute in a saturated solution at high temperature, then carefully cooling it -- such solutions are unstable. For example, a saturated solution of sodium thiosulphate ($Na_2S_2O_3$) in water at 20°C contains 20.9 g of salt per 100 $cm^3$ of water; less than this is unsaturated, and more is supersaturated at that temperature.

Solutions are also classified as dilute (relatively small amount of dissolved solute) or concentrated (relatively large amount), e.g. brine is a concentrated salt solution. Adding more solvent dilutes a solution and decreases its concentration.

### 6.3 Types of Solutions

Since solute and solvent can each be a gas, liquid, or solid, nine combinations of solutions are possible. Gas-in-gas: air, or $H_2$/He mixtures in weather balloons. Gas-in-liquid: oxygen or carbon dioxide dissolved in water. Gas-in-solid: hydrogen adsorbed on palladium. Liquid-in-gas: mist, fog, liquid air pollutants. Liquid-in-liquid: alcohol in water, benzene in toluene. Liquid-in-solid: butter, cheese. Solid-in-gas: dust particles or smoke in air. Solid-in-liquid: sugar in water. Solid-in-solid: metal alloys like brass and bronze, and opals.

### 6.4 Concentration Units

Concentration is the ratio of the amount of solute to the amount of solution (or solvent) and does not depend on the total volume or amount taken -- a sample from a bulk solution has the same concentration as the whole. Percentage concentration can be expressed four ways: %mass/mass = (mass of solute ÷ mass of solution) × 100; %mass/volume = (mass of solute in g ÷ volume of solution in $cm^3$) × 100; %volume/mass = (volume of solute in $cm^3$ ÷ mass of solution in g) × 100; %volume/volume = (volume of solute ÷ volume of solution) × 100.

Molarity (M) is the number of moles of solute dissolved in one $dm^3$ of solution -- the concentration unit most used in chemistry: Molarity = mass of solute (g) ÷ [molar mass (g/mol) × volume of solution ($dm^3$)]. A one-molar (1M) solution is prepared by dissolving one mole (the molar mass in grams) of solute in enough water to make the total volume 1 $dm^3$ -- for example, 1M NaOH is made by dissolving 40 g of NaOH and making the volume up to 1 $dm^3$. A higher molarity means a more concentrated solution.

Dilute solutions of known molarity are prepared from concentrated ones using $M_1V_1 = M_2V_2$, where $M_1/V_1$ describe the concentrated solution and $M_2/V_2$ describe the diluted one. For example, to prepare 100 $cm^3$ of 0.01M $KMnO_4$ from a 0.1M stock solution: $V_1 \times 0.1 = 0.01 \times 100$, giving $V_1 = 10$ $cm^3$ -- so 10 $cm^3$ of the 0.1M stock is measured out and diluted with water to 100 $cm^3$.

## 6.5 Solubility

Solubility is the number of grams of solute that dissolve in 100 g of solvent to form a saturated solution at a given temperature; it equals the concentration of the saturated solution. The general principle governing solubility is 'like dissolves like': ionic and polar substances (KCl, $Na_2CO_3$, $CuSO_4$, sugar, alcohol) dissolve in polar solvents like water; non-polar substances (ether, benzene, petrol) do not dissolve in water; non-polar substances (grease, paints, naphthalene) dissolve in non-polar solvents like ether or carbon tetrachloride.

Dissolution requires three events: solute particles separate from each other, solvent particles separate to make room, and solute and solvent particles attract and mix. If solute-solvent attractive forces overcome solute-solute forces, the solute dissolves; if solute-solute forces are stronger, the solute remains undissolved. When NaCl dissolves in water, the positive end of water's dipole orients toward $Cl^-$ ions and the negative end toward $Na^+$ ions -- these ion-dipole attractions are strong enough to pull the ions out of the crystal lattice.

Temperature affects solubility in three ways. When dissolving is endothermic (e.g. $KNO_3$, $NaNO_3$, KCl in water), heat is absorbed from the surroundings, the test tube feels cold, and solubility generally increases with temperature. When dissolving is exothermic (e.g. $Li_2SO_4$, $Ce_2(SO_4)_3$ in water), heat is released, the test tube feels warm, and solubility decreases with increasing temperature. For some solutes (e.g. NaCl in water), there is negligible heat change and temperature has minimal effect on solubility.

## 6.6 Comparison of Solution, Colloid, and Suspension

A true solution is a homogeneous mixture in which particles exist as individual molecules or ions (diameter about $10^{-8}$ cm); they dissolve uniformly, pass through filter paper, and do not scatter light (no Tyndall effect) -- e.g. a drop of ink fully dissolved would give a true solution, and sugar or salt solutions are classic examples.

A colloid contains larger particles (many atoms/molecules/ions per particle) that appear homogeneous but are actually a heterogeneous 'false solution' -- particles do not settle for a long time (making colloids stable) but are big enough to scatter light, producing the Tyndall effect, the key feature distinguishing colloids from true solutions. Examples include starch, albumin, soap solutions, blood, milk, ink, jelly, and toothpaste.

A suspension is a heterogeneous mixture of undissolved particles larger than $10^{-5}$ cm in diameter, big enough to see with the naked eye; these particles cannot pass through filter paper, block rather than scatter light, and settle down over time. Examples include chalk-in-water, paints, and milk of magnesia.

## Important Definitions

### Define solution.

A homogeneous mixture of two or more substances that exists as a single phase, with no visible boundaries between its components.

### Define aqueous solution.

A solution formed by dissolving a substance in water, where water is present in the greater amount and acts as the solvent.

### Define solute.

The component of a solution that is present in the smaller quantity.

### Define solvent.

The component of a solution that is present in the larger quantity and dissolves the solute(s).

### Define saturated solution.

A solution containing the maximum amount of solute that can dissolve at a given temperature, with undissolved and dissolved solute in dynamic equilibrium.

### Define unsaturated solution.

A solution that contains less solute than is required to saturate it at a given temperature, and so can dissolve more solute.

### Define supersaturated solution.

A solution more concentrated than a saturated solution, prepared by dissolving solute at high temperature and then carefully cooling the solution; it is unstable.

### Define concentration.

The proportion of solute in a solution, expressed as the ratio of the amount of solute to the amount of solution or solvent.

### Define molarity.

The number of moles of solute dissolved in one cubic decimetre ($dm^3$) of solution, denoted M.

### Define solubility.

The number of grams of solute that dissolve in 100 g of solvent to produce a saturated solution at a given temperature.

### Define colloid.

A heterogeneous 'false solution' containing particles larger than in a true solution but too small to be seen with the naked eye; it exhibits the Tyndall effect and does not settle out.

### Define suspension.

A heterogeneous mixture containing undissolved particles large enough to be seen with the naked eye, which settle down over time.

### Define the Tyndall effect.

The scattering of a beam of light by the relatively large particles present in a colloid, making the light's path visible.

### What does 'like dissolves like' mean?

The general principle that polar/ionic substances dissolve readily in polar solvents (like water), while non-polar substances dissolve in non-polar solvents (like ether or carbon tetrachloride).

### Define dilution.

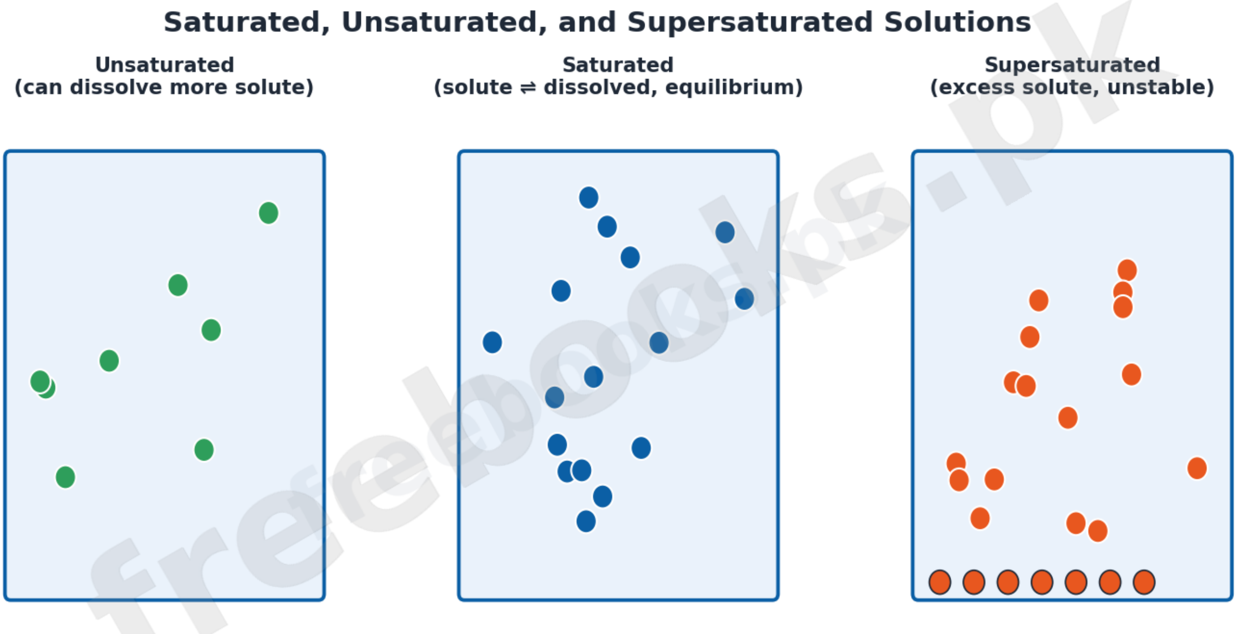
The process of adding more solvent to a solution, which decreases its concentration.

## Key Formulas

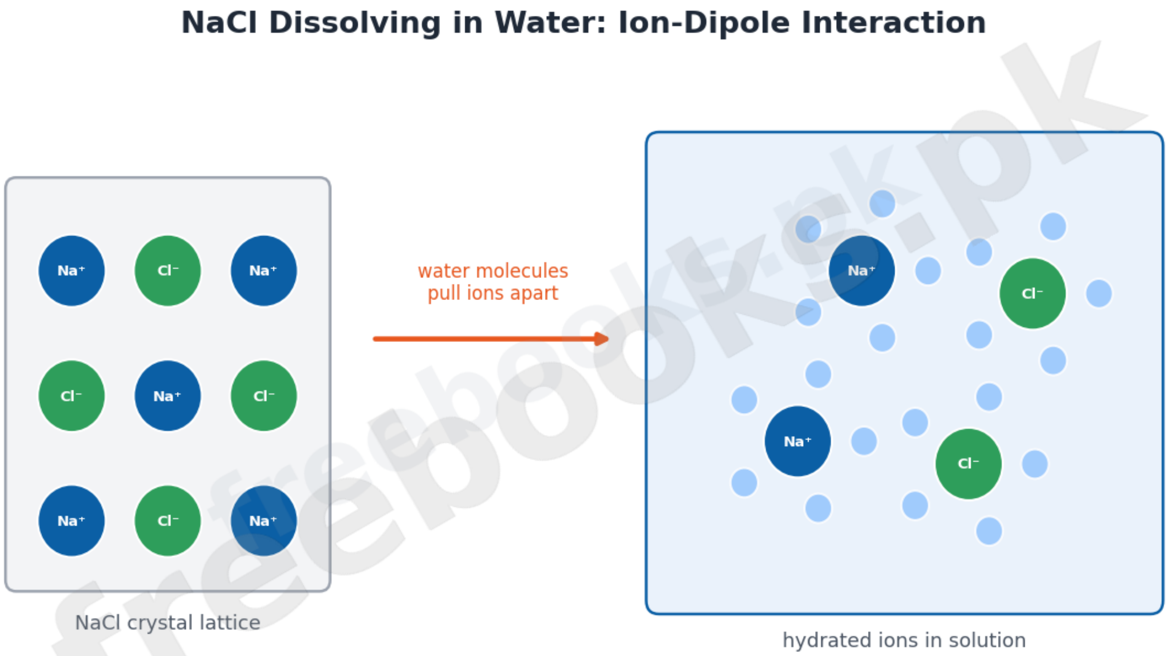
| Topic | Formula |
|---|---|
| % mass/mass | (mass of solute (g) / mass of solution (g)) × 100 |
| % mass/volume | (mass of solute (g) / volume of solution ($cm^3$)) × 100 |
| % volume/mass | (volume of solute ($cm^3$) / mass of solution (g)) × 100 |
| % volume/ volume | (volume of solute ($cm^3$) / volume of solution ($cm^3$)) × 100 |
| Molarity | M = mass of solute (g) / [molar mass (g/mol) × volume of solution ($dm^3$)] |
| Dilution formula | $M_1V_1 = M_2V_2$ (concentrated → dilute solution) |

## Diagrams

**Saturated, Unsaturated, and Supersaturated Solutions:** Visual comparison of the three saturation states, showing increasing dissolved solute and undissolved settled solute in the supersaturated case.



**NaCl Dissolving in Water: Ion-Dipole Interaction:** Illustration of the NaCl crystal lattice breaking apart as polar water molecules surround and pull away Na+ and Cl- ions to form a solution.



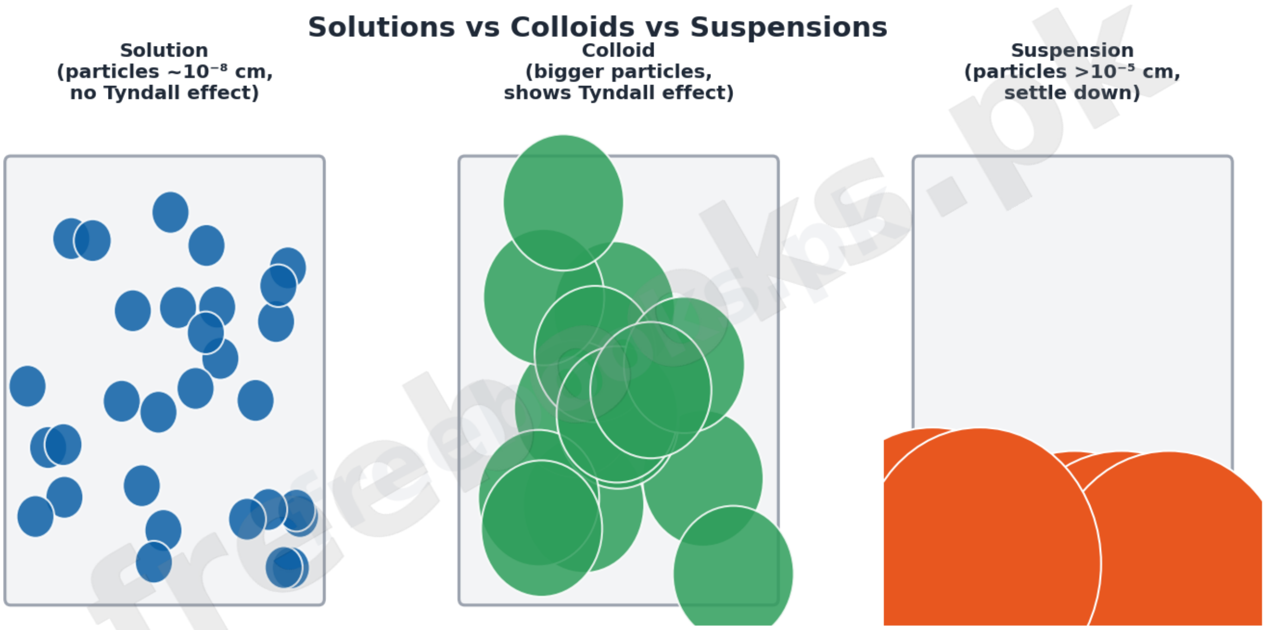

**Concentration Units at a Glance:** Summary of the five concentration units covered in this unit: %m/m, %m/v, %v/m, %v/v, and molarity, with their defining ratios.

## Concentration Units at a Glance

| | |
|---|---|
| **% mass/mass (m/m)** | g solute per 100 g solution |
| **% mass/volume (m/v)** | g solute per 100 $cm^3$ solution |
| **% volume/mass (v/m)** | $cm^3$ solute per 100 g solution |
| **% volume/volume (v/v)** | $cm^3$ solute per 100 $cm^3$ solution |
| **Molarity (M)** | moles of solute per 1 $dm^3$ solution |

*Molarity (M) = moles of solute ÷ volume of solution ($dm^3$) — the most-used unit in chemistry*

**Solutions vs Colloids vs Suspensions:** Comparison of particle size and behaviour across true solutions, colloids, and suspensions, showing why only suspensions visibly settle.



## Short Questions & Answers

### Why is a solution considered a mixture rather than a compound?

Because it shows the properties of its individual components and has a variable composition, both of which are properties of mixtures, not compounds.

### Why is water called the 'universal solvent'?

Because it dissolves the majority of compounds found in the earth's crust, more than any other single solvent.

### Why do supersaturated solutions crystallize excess solute when cooled?

Because a supersaturated solution holds more solute than is stable at the lower temperature, so the excess solute comes out of solution as crystals, leaving behind a saturated solution.

### Why does the concentration of a solution not depend on the total volume taken?

Because concentration is a ratio of solute to solution, and any sample taken from a well-mixed bulk solution has exactly the same ratio as the whole.

### Why is molarity described as the practical unit of concentration in chemistry?

Because it directly relates the number of moles of solute (the amount used in chemical reactions) to a fixed volume of solution, making stoichiometric calculations straightforward.

### Why does KNO3 make a test tube feel cold when it dissolves in water?

Because dissolving KNO3 is an endothermic process -- the salt absorbs heat from the surrounding water and test tube to break its ionic lattice, lowering the temperature.

### Why is ether, a non-polar solvent, unable to dissolve table salt?

Because table salt (NaCl) is an ionic solid, and ionic substances only dissolve well in polar solvents whose molecules can surround and stabilize the separated ions -- non-polar ether cannot do this.

### Why do colloids not settle down like suspensions?

Because colloidal particles, although larger than those in a true solution, are still small and light enough that random molecular collisions (Brownian motion) keep them dispersed rather than settling under gravity.

**Why can suspension particles be removed by filtration but solution particles cannot?**
Because suspension particles are large enough to be physically trapped by the pores of filter paper, while solute particles in a true solution are the size of individual molecules or ions and pass straight through.

**Why does milk show the Tyndall effect while sugar solution does not?**
Because milk is a colloid with particles large enough to scatter a beam of light, while sugar solution is a true solution whose particles are too small to scatter light at all.

## Long Questions & Answers

### Explain the concept of saturated, unsaturated, and supersaturated solutions, including how a supersaturated solution is prepared.

#### What is a saturated solution?

A saturated solution is one that contains the maximum amount of solute that can dissolve in the solvent at a given temperature. At the particle level, it represents a dynamic equilibrium in which undissolved solute continues to dissolve while dissolved solute continues to crystallize out, keeping the net dissolved amount constant.

#### What is an unsaturated solution?

An unsaturated solution contains less solute than the amount required to saturate it at that temperature. Because it has not reached its maximum capacity, an unsaturated solution can still dissolve additional solute until it becomes saturated.

#### What is a supersaturated solution, and why is it unstable?

A supersaturated solution contains more dissolved solute than a saturated solution normally would at that temperature, making it more concentrated. It is unstable because this extra solute is only held in solution under special conditions; any disturbance, such as adding a seed crystal or a temperature change, can cause the excess solute to rapidly crystallize out.

#### How is a supersaturated solution prepared in practice?

A saturated solution is first prepared at a high temperature, where more solute can dissolve. This hot saturated solution is then carefully and slowly cooled to a lower temperature; because solubility usually decreases with cooling, the solution now holds more solute than is stable at the lower temperature, making it supersaturated, until the excess eventually crystallizes out.

### Explain how solute-solvent interaction determines whether a solute dissolves, using NaCl in water as an example.

#### What three events must occur for a solute to dissolve in a solvent?

First, solute particles must separate from each other. Second, solvent particles must separate to make room for the incoming solute particles. Third, the solute and solvent particles must attract one another and mix together to form a uniform solution.

#### What determines whether these events actually happen?

Dissolution depends on the relative strength of three types of attractive forces: solute-solute forces, solvent-solvent forces, and solute-solvent forces. If the new solute-solvent forces are strong enough to overcome the solute-solute forces holding the solid together, the solute dissolves; if the solute-solute forces are stronger, the solute remains undissolved.

#### How does this apply specifically to NaCl dissolving in water?

In solid NaCl, strong ionic attractions hold $Na^+$ and $Cl^-$ ions in a rigid crystal lattice. When NaCl is added to water, the polar water molecules orient themselves so their negative (oxygen) end faces the $Na^+$ ions and their positive (hydrogen) end faces the $Cl^-$ ions.

#### What is the outcome of these ion-dipole attractions?

These ion-dipole attractions between the water molecules and the $Na^+$ and $Cl^-$ ions are strong enough to overcome the ionic forces holding the crystal together, pulling the ions out of their lattice positions one by one. The separated ions become surrounded by water molecules, and the solid NaCl fully dissolves into solution.

## Multiple Choice Questions (MCQs)

**Mist is an example of which type of solution? (A) liquid in gas (B) gas in liquid (C) solid in gas (D) gas in solid**

Correct answer: (A) liquid in gas. Mist consists of tiny liquid water droplets dispersed in air (a gas), making it a liquid-in-gas solution.

**Which one of the following is a 'liquid in solid' solution? (A) sugar in water (B) butter (C) opal (D) fog**

Correct answer: (B) butter. Butter is a liquid (fat/water droplets) dispersed within a solid matrix, making it a liquid-in-solid solution.

**Concentration is a ratio of: (A) solvent to solute (B) solute to solution (C) solvent to solution (D) both solute-to-solution and solvent-to-solution**

Correct answer: (B) solute to solution. Concentration is defined as the ratio of the amount of solute to the amount of solution (or sometimes to the solvent).

**Which one of the following solutions contains the most water (is most dilute) among equal volumes? (A) 2 M (B) 1 M (C) 0.5 M (D) 0.25 M**

Correct answer: (D) 0.25 M. A 0.25 M solution has the least solute per unit volume among these choices, meaning it contains proportionally the most water/solvent.

**A 5 percent (m/m) sugar solution means: (A) 5 g of sugar is dissolved in 90 g of water (B) 5 g of sugar is dissolved in 100 g of water (C) 5 g of sugar is dissolved in 105 g of water (D) 5 g of sugar is dissolved in 95 g of water**

Correct answer: (D) 5 g of sugar is dissolved in 95 g of water. In %m/m, the mass of solute plus mass of solvent equals 100 g total, so 5 g sugar must be dissolved in 95 g of water to make 100 g of solution.

**If solute-solute forces are stronger than solute-solvent forces, the solute: (A) dissolves readily (B) does not dissolve (C) dissolves slowly (D) dissolves and precipitates**

Correct answer: (B) does not dissolve. When the solute's own internal attractive forces are stronger than the attraction offered by the solvent, the solvent cannot pull the solute apart, so it does not dissolve.

**Which of the following will show negligible effect of temperature on its solubility? (A) KCl (B) $KNO_3$ (C) $NaNO_3$ (D) NaCl**

Correct answer: (D) NaCl. Dissolving NaCl in water involves very little heat change, so its solubility is only minimally affected by temperature changes.

**Which one of the following is a heterogeneous mixture? (A) milk (B) ink (C) milk of magnesia (D) sugar solution**

Correct answer: (C) milk of magnesia. Milk of magnesia is a suspension of solid magnesium hydroxide particles in water, which is a heterogeneous mixture that settles over time, unlike milk and ink which behave as stable colloids here.

**The Tyndall effect is caused by: (A) blockage of the beam of light (B) non-scattering of the beam of light (C) scattering of the beam of light (D) the beam of light passing straight through**

Correct answer: (C) scattering of the beam of light. The Tyndall effect occurs because colloidal particles are large enough to scatter light in many directions as the beam passes through, making its path visible.

**Molarity is the number of moles of solute dissolved in: (A) 1 kg of solution (B) 100 g of solvent (C) 1 $dm^3$ of solvent (D) 1 $dm^3$ of solution**

Correct answer: (D) 1 $dm^3$ of solution. By definition, molarity (M) is the number of moles of solute per 1 $dm^3$ (one litre) of the total solution, not the solvent alone.

## Quick Revision Summary

- Solution = homogeneous mixture; solute = smaller quantity; solvent = larger quantity
- Aqueous solution = water is the solvent; water is the 'universal solvent'
- Saturated: max solute at equilibrium; unsaturated: can dissolve more; supersaturated: unstable excess
- 9 solution types from combinations of gas/liquid/solid solute + solvent
- 4 percentage concentration types: %m/m, %m/v, %v/m, %v/v -- know which quantity is 100 in each

- Molarity M = moles of solute ÷ volume of solution ($dm^3$); practical unit in chemistry
- Dilution formula: $M_1V_1 = M_2V_2$ (concentrated → dilute)
- 'Like dissolves like': polar/ionic in polar solvents, non-polar in non-polar solvents
- Temperature effect on solubility: endothermic ($KNO_3$, NaCl-ish neutral) increases with heat; exothermic ($Li_2SO_4$) decreases
- Solution (no Tyndall, ~$10^{-8}$ cm) < Colloid (Tyndall effect, doesn't settle) < Suspension (settles, visible particles)

## Exam Tips

- Memorize all four percentage formulas precisely -- exams often test which quantity (solute/solution/solvent) forms the denominator in each
- Practice molarity numericals repeatedly: converting $cm^3$ to $dm^3$ is the most common calculation mistake
- Learn the dilution formula $M_1V_1=M_2V_2$ and practice identifying which values are 'concentrated' (1) vs 'dilute' (2)
- Be ready to classify a given example (mist, butter, brass, etc.) into one of the nine solution types
- Know at least 2-3 examples each of endothermic and exothermic dissolution for the temperature-solubility questions
- Practice distinguishing solution vs colloid vs suspension using the Tyndall effect and settling behaviour as the key tests
- Remember 'like dissolves like' with the specific example pairs given (NaCl in water; grease in $CCl_4$/ether)

# Unit 7: Electrochemistry

Electrochemistry is the branch of chemistry dealing with the relationship between electricity and chemical reactions, centered on oxidation-reduction (redox) reactions. Redox reactions either occur spontaneously and produce electricity (as in galvanic cells) or are driven by electricity to force a non-spontaneous reaction to occur (as in electrolytic cells).

This unit covers the three complementary definitions of oxidation and reduction; oxidation states and the rules for assigning them; oxidizing and reducing agents; redox reactions with oxidation-number tracking; the construction and working of electrolytic cells and galvanic (Daniel) cells; industrial electrochemical processes including the Downs cell (sodium metal) and Nelson's cell (NaOH); and corrosion, rusting, and its prevention through methods including galvanizing, tin coating, and electroplating.

## Learning Objectives

- Define oxidation and reduction in terms of gain/loss of oxygen, hydrogen, and electrons
- Identify oxidizing and reducing agents in a redox reaction
- Define oxidation state and apply the rules for assigning oxidation numbers
- Determine the oxidation number of an atom of any element in a compound or ion
- Sketch and label an electrolytic cell, identifying the cathode, anode, and ion movement
- Sketch and label a Daniel (galvanic) cell, showing electron flow and half-cell reactions
- Distinguish between electrolytic and galvanic (voltaic) cells
- Describe the industrial manufacture of sodium metal (Downs cell) and sodium hydroxide (Nelson's cell)
- Explain electrolytic refining of copper
- Define corrosion and describe the electrochemical mechanism of rusting
- Summarize methods used to prevent corrosion, including electroplating

## Key Concepts

### 7.1 Oxidation and Reduction Reactions

Oxidation and reduction can first be understood in terms of oxygen/hydrogen transfer: oxidation is the addition of oxygen or removal of hydrogen, while reduction is the addition of hydrogen or removal of oxygen. Both processes occur simultaneously in a reaction -- where there is oxidation, there is reduction. For example, in the reaction between zinc oxide and carbon, oxygen is removed from zinc oxide (reduction) and added to carbon (oxidation); in the reaction between hydrogen sulphide and chlorine, hydrogen is removed from $H_2S$ (oxidation) and added to chlorine (reduction).

Many redox reactions do not involve oxygen or hydrogen at all, so a more general definition uses electron transfer: oxidation is the loss of electrons by an atom or ion (e.g. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$), while reduction is the gain of electrons (e.g. $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$). A reaction between sodium and chlorine illustrates this: sodium loses an electron (oxidation) to become $Na^+$, while chlorine simultaneously gains that electron (reduction) to become $Cl^-$, and the two ions then combine to form NaCl.

### 7.2 Oxidation State and Rules for Assigning Oxidation Numbers

Oxidation state (or oxidation number, O.N.) is the apparent charge assigned to an atom of an element in a molecule or ion -- for example, in HCl, H has O.N. +1 and Cl has O.N. -1. The key rules are: the O.N. of any element in its free (uncombined) state is zero; the O.N. of a monoatomic ion equals its charge; Group 1 elements are always +1, Group 2 always +2, and Group 13 always +3; hydrogen is +1 in most compounds but -1 in metal hydrides; oxygen is -2 in most compounds, -1 in peroxides, and +2 in $OF_2$; the more electronegative atom in a substance takes the negative oxidation number; in neutral molecules the sum of all oxidation numbers is zero; and in ions, the sum of oxidation numbers equals the ion's overall charge.

When assigning oxidation numbers, the sign is written before the number (e.g. +2), whereas valency (the apparent charge on an atom/ion) is written with the number first and the sign after (e.g. 2+). For example, in $HNO_3$, using H = +1 and O = -2: [+1] + [O.N. of N] + 3[-2] = 0, giving O.N. of N = +5. Similarly, in $H_2SO_4$, the oxidation number of sulphur works out to +6, and in $KClO_3$, the oxidation number of chlorine works out to +5.

### 7.3 Oxidizing and Reducing Agents

An oxidizing agent is a species that oxidizes another substance by taking electrons from it -- in doing so, the oxidizing agent itself is reduced (gains electrons). Non-metals, being more electronegative, tend to act as oxidizing agents. A reducing agent is a species that reduces another substance by donating electrons to it -- the reducing agent itself is oxidized (loses electrons) in the process. Almost all metals are good reducing agents because they readily lose electrons.

### 7.4 Oxidation-Reduction Reactions

A redox reaction is one in which the oxidation state of one or more substances changes. Tracking oxidation numbers through a reaction reveals which species are oxidized and which are reduced. For example, in $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$, zinc's oxidation number rises from 0 to +2 (oxidized) while hydrogen's falls from +1 to 0 (reduced). Similarly, in $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, hydrogen is oxidized from 0 to +1 while oxygen is reduced from 0 to -2.

### 7.5 Electrochemical Cells

An electrochemical cell contains two electrodes dipped in an electrolyte solution, either driving a non-spontaneous reaction with electric current (electrolytic cell) or generating electric current from a spontaneous reaction (galvanic/voltaic cell). Electrolytes are substances that conduct electricity in aqueous solution or molten state. Strong electrolytes (e.g. NaCl, NaOH, $H_2SO_4$) ionize almost completely; weak electrolytes (e.g. $CH_3COOH$, $Ca(OH)_2$) ionize only slightly and conduct poorly; non-electrolytes (e.g. sugar solution, benzene) do not ionize at all and do not conduct.

An electrolytic cell uses electric current to drive a non-spontaneous reaction (electrolysis) -- the anode carries a positive charge and the cathode a negative charge. Anions migrate to the anode and are oxidized there (lose electrons); cations migrate to the cathode and are reduced there (gain electrons). For fused

NaCl electrolysis: anode (oxidation): $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$; cathode (reduction): $2Na^+(l) + 2e^- \rightarrow 2Na(l)$. Pure water is a very weak electrolyte; adding a little acid improves conductivity, and electrolysis of acidified water produces $O_2$ at the anode and $H_2$ at the cathode.

A galvanic (voltaic) cell, such as the Daniel cell, uses a spontaneous redox reaction to generate electric current -- here the anode is negative and the cathode is positive (the reverse of an electrolytic cell). A Daniel cell has two half-cells joined by a salt bridge: a zinc electrode in $ZnSO_4$ solution and a copper electrode in $CuSO_4$ solution. Zinc loses electrons more readily than copper, so oxidation ($Zn \rightarrow Zn^{2+} + 2e^-$) occurs at the zinc electrode, electrons flow through the external wire to the copper electrode, and reduction ($Cu^{2+} + 2e^- \rightarrow Cu$) occurs there. The overall reaction $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$ produces electric current, the same principle used in everyday batteries.

### 7.6 Electrochemical Industries

Sodium metal is manufactured industrially by electrolysis of fused NaCl in the Downs cell, a circular furnace with a central graphite anode and a surrounding iron cathode, separated by steel gauze to prevent product mixing. $Cl^-$ ions are oxidized to $Cl_2$ gas at the anode; $Na^+$ ions are reduced to molten sodium at the cathode, which floats on the denser molten salt and is collected separately. Overall: $2NaCl(fused) \rightarrow Cl_2(g) + 2Na(l)$.

Sodium hydroxide (caustic soda) is manufactured industrially by electrolysis of brine (aqueous NaCl) in Nelson's cell, which has a graphite anode suspended inside a U-shaped perforated iron cathode lined with an asbestos diaphragm. $Cl^-$ ions are discharged at the anode producing $Cl_2$ gas; $H^+$ ions are discharged at the cathode producing $H_2$ gas, while NaOH solution percolates out and collects in a catch basin. Overall: $2NaCl(aq) + 2H_2O(l) \rightarrow H_2(g) + Cl_2(g) + 2NaOH(aq)$.

Impure copper is purified by electrolytic refining: impure copper serves as the anode and a pure copper plate as the cathode, in a copper sulphate electrolyte. At the anode, copper atoms dissolve as $Cu^{2+}$ ions (oxidation); at the cathode, $Cu^{2+}$ ions gain electrons and deposit as pure copper (reduction), leaving impurities behind.

### 7.7 Corrosion and Its Prevention

Corrosion is the slow, continuous eating away of a metal by its surrounding medium through a redox reaction; the corrosion of iron specifically is called rusting, which requires moist air (both water vapour and oxygen must be present). At the anodic region (stains/dents on the iron surface), Fe is oxidized: $2Fe(s) \rightarrow 2Fe^{2+}(aq) + 4e^-$. Electrons flow through the iron to a cathodic region of high $O_2$ concentration, where oxygen is reduced in the presence of $H^+$ ions (supplied by carbonic acid from dissolved $CO_2$): $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$. The $Fe^{2+}$ formed later reacts with more oxygen and water to form rust, $Fe_2O_3 \cdot nH_2O$, which is porous and does not protect the metal beneath, so rusting continues until the iron is consumed.

Corrosion prevention methods include: removing surface stains (which act as corrosion sites); applying paints or grease as protective, weatherproof barriers; alloying, such as combining iron with chromium and nickel to form stainless steel; and metallic coating, where corrosion-resistant metals like zinc, tin, or chromium are applied over iron. Galvanizing (zinc coating) protects iron even after the coating is broken, because zinc corrodes preferentially; tin coating protects only while the tin layer stays intact, since once broken, a galvanic cell forms and the exposed iron rusts rapidly.

Electroplating deposits one metal over another using electrolysis: the anode is made of the metal to be deposited, the cathode is the object to be plated, and the electrolyte is a salt solution of the depositing metal. In silver plating, a silver strip anode dissolves ($Ag \rightarrow Ag^+ + e^-$) and $Ag^+$ ions migrate to and deposit on the cathode object ($Ag^+ + e^- \rightarrow Ag$). Chromium plating uses an antimonial lead anode with $Cr_2(SO_4)_3$ electrolyte, and steel is usually first plated with nickel or copper (for better adhesion) before the final chromium layer, which resists corrosion and gives a bright, silvery finish.

## Important Definitions

### Define oxidation (electron concept).

The loss of electrons by an atom or ion during a chemical reaction.

### Define reduction (electron concept).

The gain of electrons by an atom or ion during a chemical reaction.

### Define oxidation state (oxidation number).

The apparent charge assigned to an atom of an element in a molecule or ion.

### Define oxidizing agent.

A species that oxidizes another substance by removing electrons from it, and is itself reduced in the process.

### Define reducing agent.

A species that reduces another substance by donating electrons to it, and is itself oxidized in the process.

### Define redox reaction.

A chemical reaction in which the oxidation states of one or more substances change, involving simultaneous oxidation and reduction.

### Define electrolyte.

A substance that can conduct electricity in its aqueous solution or molten state.

### Define electrolytic cell.

An electrochemical cell in which a non-spontaneous chemical reaction is driven by passing electric current through an electrolyte.

### Define galvanic (voltaic) cell.

An electrochemical cell in which a spontaneous chemical reaction takes place and generates electric current.

### Define electrolysis.

The chemical decomposition of a compound into its components by passing an electric current through its solution or molten state.

### Define corrosion.

The slow and continuous eating away of a metal by the action of its surrounding medium, occurring through a redox reaction.

### Define rusting.

The specific term for the corrosion of iron, forming hydrated iron(III) oxide (Fe2O3·nH2O) in the presence of moist air.

### Define electroplating.

The process of depositing one metal over another by means of electrolysis, to protect against corrosion or improve appearance.

### Define galvanizing.

The process of coating a thin protective layer of zinc onto iron or steel to prevent corrosion.

### Define salt bridge.

A U-shaped tube containing a saturated electrolyte solution that connects the two half-cells of a galvanic cell, keeping both solutions electrically neutral by allowing ion migration.

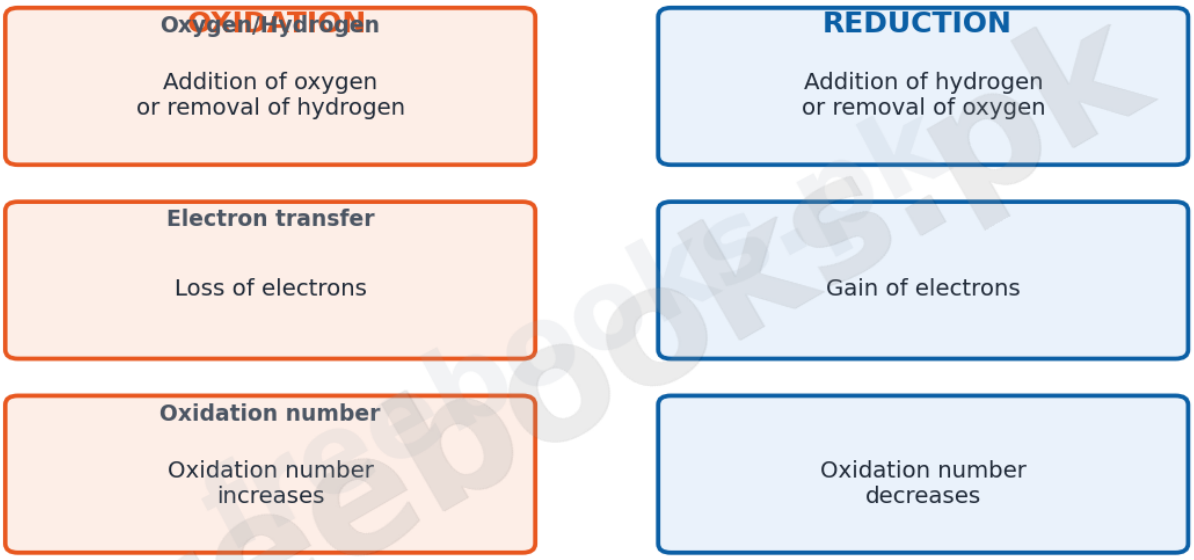
## Key Formulas

| Topic | Formula |
|---|---|
| Sum of O.N. in a neutral molecule | Σ(oxidation numbers) = 0 |
| Sum of O.N. in an ion | Σ(oxidation numbers) = charge on the ion |
| Oxidation (electron loss) | $M(s) \rightarrow M^{n+}(aq) + n\,e^-$ |
| Reduction (electron gain) | $X + n\,e^- \rightarrow X^{n-}$ |
| Downs cell overall reaction | $2NaCl(fused) \rightarrow Cl_2(g) + 2Na(l)$ |
| Nelson's cell overall reaction | $2NaCl(aq) + 2H_2O(l) \rightarrow H_2(g) + Cl_2(g) + 2NaOH(aq)$ |
| Overall rusting reaction | $2Fe(s) + O_2(g) + 4H^+(aq) \rightarrow 2Fe^{2+}(aq) + 2H_2O(l); \rightarrow Fe_2O_3{\cdot}nH_2O$ (rust) |

## Diagrams

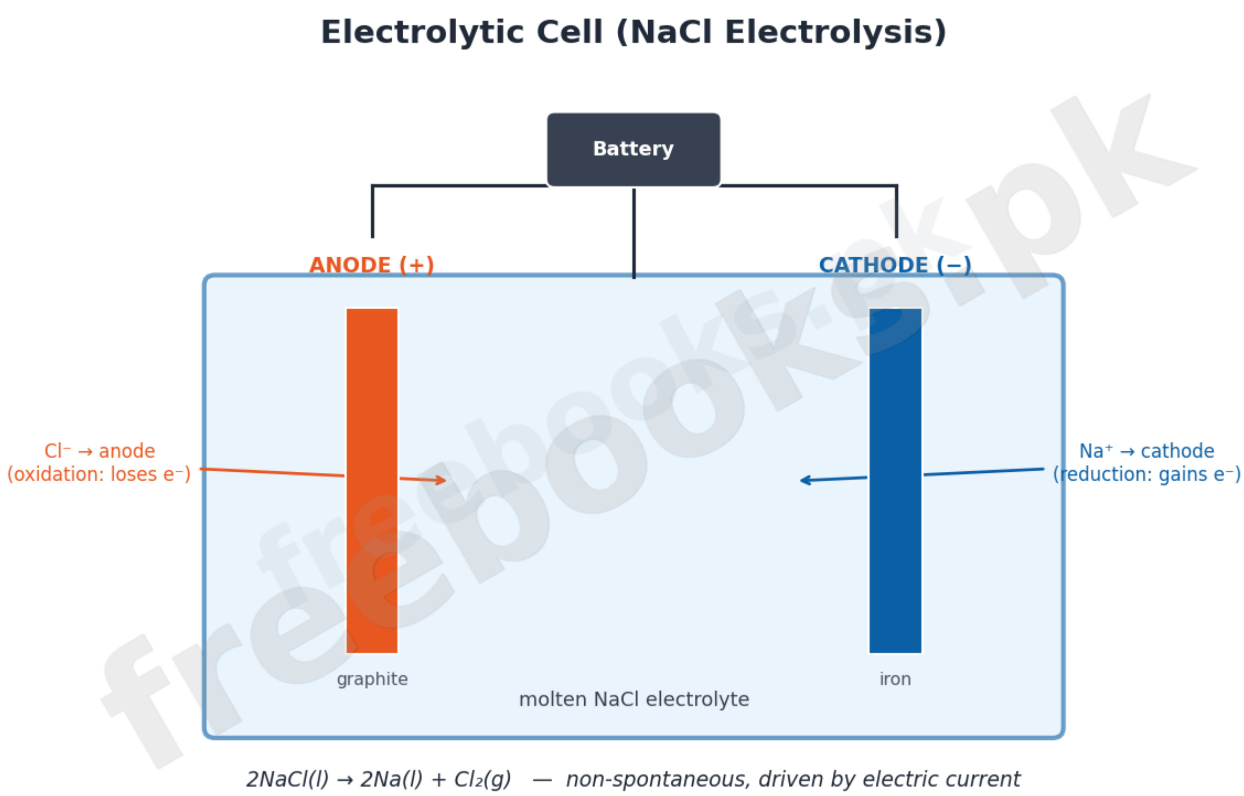
**Oxidation vs Reduction: Three Ways to Define Each:** Side-by-side summary of the oxygen/hydrogen, electron-transfer, and oxidation-number definitions of oxidation and reduction.

**Oxidation vs Reduction: Three Ways to Define Each**

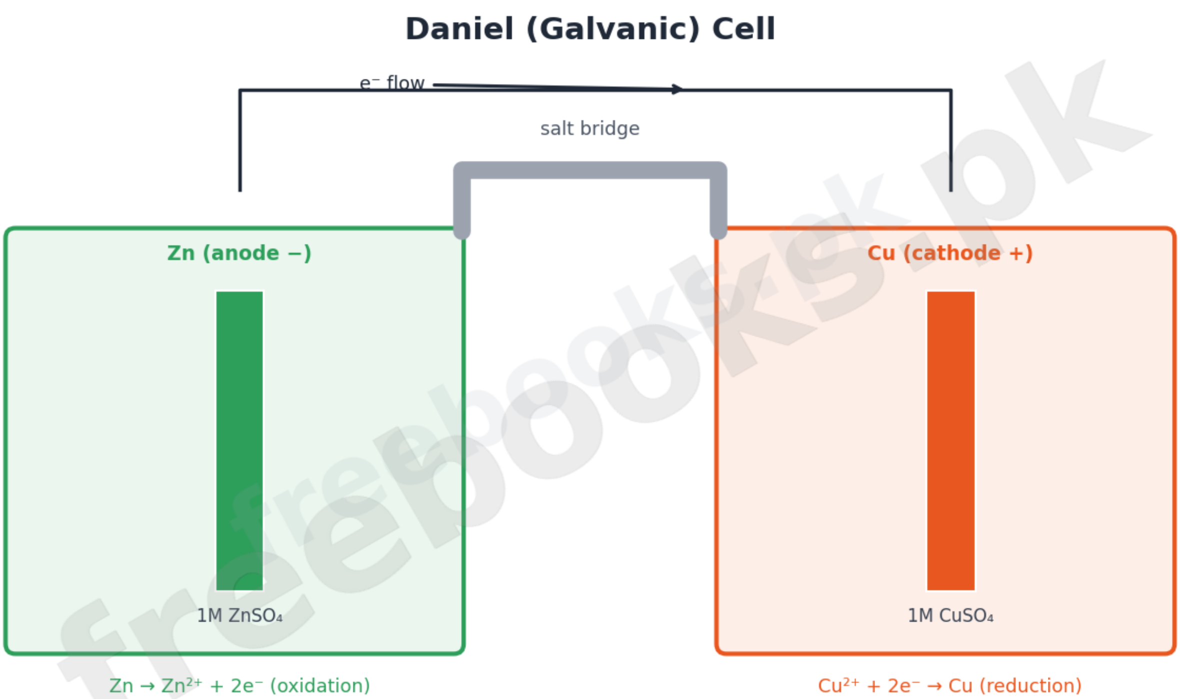



*Where there is oxidation, there is reduction — both occur simultaneously in a redox reaction.*

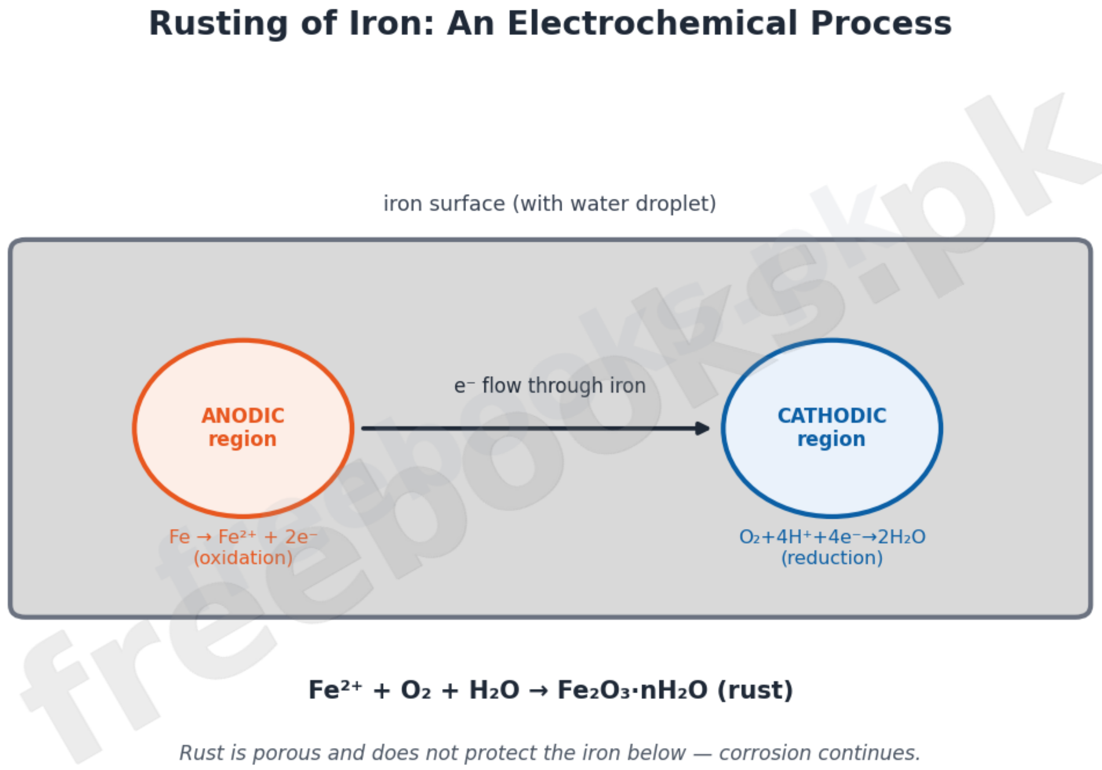
**Electrolytic Cell (NaCl Electrolysis):** Labelled diagram of an electrolytic cell showing the battery, anode, cathode, and ion migration during electrolysis of molten NaCl.




**Daniel (Galvanic) Cell:** Labelled diagram of a Daniel cell showing the zinc and copper half-cells, salt bridge, external wire, and direction of electron flow.

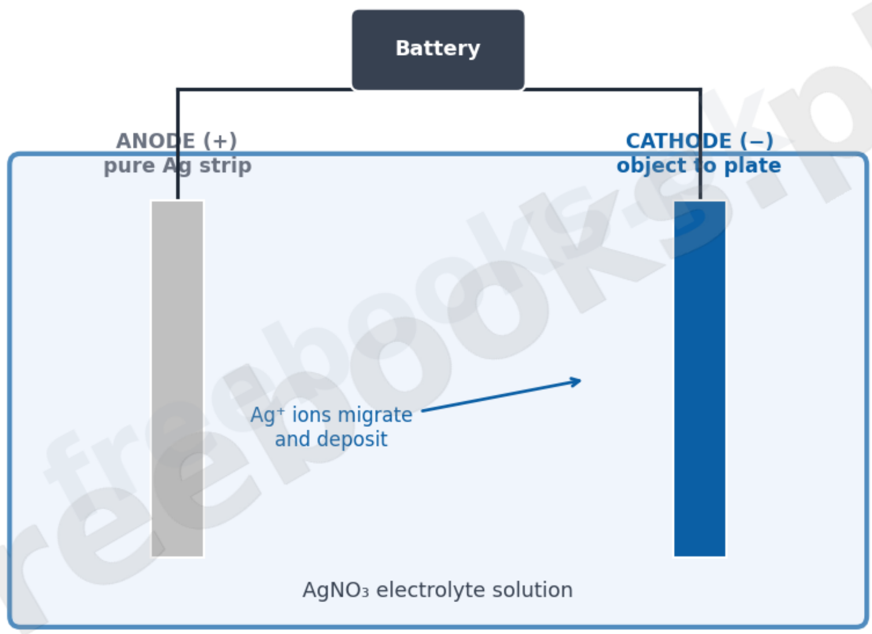


**Rusting of Iron: An Electrochemical Process:** Diagram showing the anodic (oxidation) and cathodic (reduction) regions on an iron surface that together produce rust.




**Electroplating Setup:** Diagram of an electroplating cell (e.g. silver plating) showing the anode, cathode (object), electrolyte, and ion deposition.

**Electroplating Setup (e.g. Silver Plating)**




*Anode: $Ag \rightarrow Ag^+ + e^-$     Cathode: $Ag^+ + e^- \rightarrow Ag$ (deposits on object)*

## Short Questions & Answers

### Why is the reaction between magnesium and oxygen (2Mg + O2 → 2MgO) considered a redox reaction even though it only shows addition of oxygen?

Because while oxygen is being added to magnesium (oxidation), magnesium is simultaneously being reduced in the sense that it loses electrons to oxygen, and oxygen itself is reduced by gaining those electrons -- so both oxidation and reduction occur together, satisfying the definition of a redox reaction.

### Why must the sign precede the number when writing an oxidation state, but follow the number for valency?

This is simply the established chemical convention: oxidation numbers are written as +2 or -2, while valency (the apparent ionic charge) is written as 2+ or 2-, allowing the two related but distinct concepts to be distinguished at a glance.

### Why are non-metals generally good oxidizing agents?

Because non-metals are highly electronegative and have a strong tendency to attract and gain electrons, which means they readily take electrons from other substances (oxidizing them) while being reduced themselves.

### Why does the anode carry a positive charge in an electrolytic cell but a negative charge in a galvanic cell?

In an electrolytic cell, the anode is connected to the positive terminal of an external battery, giving it a positive charge. In a galvanic cell, the anode is the electrode where oxidation spontaneously occurs and electrons are released, making it electron-rich and therefore negatively charged relative to the cathode.

### Why is a salt bridge necessary in a Daniel cell?

Because as the reaction proceeds, positive ions build up in the anode half-cell and negative ions build up in the cathode half-cell; the salt bridge allows ions to migrate between the two half-cells to keep both solutions electrically neutral, allowing the reaction to continue.

### Why does rusting require both moisture and oxygen to occur?

Because rusting is an electrochemical redox process: oxygen is needed at the cathodic region to be reduced (gaining the electrons released at the anode), while water provides the medium through which ions can migrate between the anodic and cathodic regions -- without both, the redox circuit cannot be completed.

### Why does acidic water accelerate the rusting of iron?

Because the reduction half-reaction at the cathodic region requires H+ ions to reduce oxygen to water, and an acidic medium supplies these H+ ions in greater abundance, speeding up the overall rusting process.

### Why does galvanizing protect iron even after the zinc coating is scratched or broken?

Because zinc is more reactive than iron and corrodes preferentially, acting as a sacrificial anode -- as long as some zinc remains in contact with the iron, it will continue to be oxidized instead of the iron, even where the coating is damaged.

### Why does tin-plated iron rust rapidly once the tin layer is broken?

Because tin is less reactive than iron, so once the protective tin layer is broken and iron is exposed, a galvanic cell forms in which the iron (more reactive) becomes the anode and corrodes rapidly, actually accelerating rusting at the exposed point.

### Why is nickel or copper often plated onto steel before the final chromium layer in electroplating?

Because chromium does not adhere well directly to steel and would allow moisture to seep through and strip the metal off; a nickel or copper underlayer provides good adhesion, allowing the chromium plated over it to last much longer.

## Long Questions & Answers

### Explain the concept of oxidation and reduction using the three complementary definitions, with suitable examples.

#### How are oxidation and reduction defined in terms of oxygen and hydrogen?

Oxidation is the addition of oxygen or the removal of hydrogen from a substance during a reaction, while reduction is the addition of hydrogen or the removal of oxygen. For example, in the reaction between zinc oxide and carbon, oxygen is removed from zinc oxide (reduction) while it is added to carbon (oxidation).

#### How are oxidation and reduction defined in terms of electron transfer?

Oxidation is the loss of electrons by an atom or ion, such as $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$, while reduction is the gain of electrons, such as $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$. This definition is more general because it covers redox reactions that do not involve oxygen or hydrogen at all.

#### How are oxidation and reduction defined in terms of oxidation number?

Oxidation corresponds to an increase in the oxidation number of an element, while reduction corresponds to a decrease. For example, when Zn(0) becomes $Zn^{2+}$, its oxidation number rises from 0 to +2, confirming oxidation; when $Cl_2(0)$ becomes $Cl^-$, chlorine's oxidation number falls from 0 to -1, confirming reduction.

#### Why must oxidation and reduction always occur together?

Because electrons that are lost by one species during oxidation must be gained by another species, which undergoes reduction -- the two processes are two halves of the same electron-transfer event, so a redox reaction always contains both simultaneously, even if only one type of change (like oxygen addition) is immediately visible.

## Describe the construction and working of a Daniel cell, and explain how it differs from an electrolytic cell.

### How is a Daniel cell constructed?

A Daniel cell consists of two half-cells connected by a salt bridge: the left half-cell has a zinc electrode dipped in 1M zinc sulphate solution, and the right half-cell has a copper electrode dipped in 1M copper sulphate solution. Each electrode is connected through a wire to an external circuit.

### How does the Daniel cell generate electric current?

Zinc has a greater tendency to lose electrons than copper, so oxidation occurs at the zinc electrode ($Zn \rightarrow Zn^{2+} + 2e^-$), releasing electrons that flow through the external wire to the copper electrode. At the copper electrode, these electrons are gained by copper ions in solution, which are reduced and deposit as copper metal ($Cu^{2+} + 2e^- \rightarrow Cu$), and this flow of electrons through the wire constitutes the electric current.

### What is the role of the salt bridge?

The salt bridge is a U-shaped tube containing a saturated electrolyte in a jelly-like support, sealed at both ends with porous material. It allows ions to migrate between the two half-cells, keeping both solutions electrically neutral as the reaction proceeds, which allows the redox reaction to continue producing current.

### How does a Daniel cell (galvanic cell) fundamentally differ from an electrolytic cell?

A galvanic cell like the Daniel cell uses a spontaneous redox reaction to generate electric current, converting chemical energy into electrical energy, with the anode being negative and the cathode positive. An electrolytic cell does the opposite: it uses an external electric current to force a non-spontaneous reaction to occur, converting electrical energy into chemical energy, with the anode being positive and the cathode negative.

## Multiple Choice Questions (MCQs)

### Spontaneous chemical reactions take place in: (A) electrolytic cell (B) galvanic cell (C) Nelson's cell (D) Downs cell

Correct answer: (B) galvanic cell. A galvanic (voltaic) cell is defined as one in which a spontaneous redox reaction occurs and generates electric current, unlike electrolytic cells which require an external current to drive a non-spontaneous reaction.

### Formation of water from hydrogen and oxygen ($2H_2 + O_2 \rightarrow 2H_2O$) is a: (A) redox reaction (B) acid-base reaction (C) neutralization reaction (D) decomposition reaction

Correct answer: (A) redox reaction. Hydrogen is oxidized (0 to +1) and oxygen is reduced (0 to -2) in this reaction, making it a classic redox reaction.

### Which one of the following is NOT an electrolytic cell? (A) Downs cell (B) galvanic cell (C) Nelson's cell (D) both galvanic cell and Nelson's cell

Correct answer: (B) galvanic cell. A galvanic cell is fundamentally different from an electrolytic cell -- it generates current from a spontaneous reaction rather than using current to force a non-spontaneous one, so it is not an electrolytic cell (unlike Downs and Nelson's cells, which are both electrolytic).

### The oxidation number of chromium in $K_2Cr_2O_7$ is: (A) +2 (B) +6 (C) +7 (D) +14

Correct answer: (B) +6. Using O.N. of K = +1 and O = -2: 2(+1) + 2(O.N. Cr) + 7(-2) = 0, giving 2 + 2(O.N. Cr) - 14 = 0, so O.N. of Cr = +6.

### Which one of the following is NOT an electrolyte? (A) sugar solution (B) sulphuric acid solution (C) lime solution (D) sodium chloride solution

Correct answer: (A) sugar solution. Sugar solution is a non-electrolyte because sugar molecules do not ionize in water and therefore cannot conduct electricity.

### The most common example of corrosion is: (A) chemical decay (B) rusting of iron (C) rusting of aluminium (D) rusting of tin

Correct answer: (B) rusting of iron. Rusting of iron is explicitly identified as the most common and well-known example of the general phenomenon of corrosion.

### In Nelson's cell, which gas is produced at the cathode? (A) $Cl_2$ (B) $H_2$ (C) $O_3$ (D) $O_2$

Correct answer: (B) $H_2$. At the cathode of Nelson's cell, $H^+$ ions are discharged and reduced, producing hydrogen gas ($H_2$), while chlorine gas forms at the anode.

**During the formation of water from hydrogen and oxygen, which of the following does NOT occur? (A) hydrogen is oxidized (B) oxygen is reduced (C) oxygen gains electrons (D) hydrogen behaves as an oxidizing agent**

Correct answer: (D) hydrogen behaves as an oxidizing agent. Hydrogen is oxidized (loses control of electrons) in this reaction and therefore acts as a reducing agent, not an oxidizing agent -- oxygen is the oxidizing agent here since it gains electrons and is reduced.

**The chemical formula of rust is: (A) $Fe_2O_3 \cdot nH_2O$ (B) $Fe_2O_3$ (C) $Fe(OH)_3 \cdot nH_2O$ (D) $Fe(OH)_3$**

Correct answer: (A) $Fe_2O_3 \cdot nH_2O$. Rust is hydrated iron(III) oxide, correctly written as $Fe_2O_3 \cdot nH_2O$, reflecting its variable water content.

**In the redox reaction between Zn and HCl, the oxidizing agent is: (A) Zn (B) $H^+$ (C) $Cl^-$ (D) $H_2$**

Correct answer: (B) $H^+$. $H^+$ ions gain electrons and are reduced to $H_2$ gas in this reaction, making $H^+$ the species that gets reduced -- and the species that is reduced is, by definition, the oxidizing agent.

## Quick Revision Summary

- Oxidation = add O / remove H / lose $e^-$ / O.N. increases; Reduction = add H / remove O / gain $e^-$ / O.N. decreases
- Oxidizing agent = itself reduced (gains $e^-$), usually a non-metal; reducing agent = itself oxidized (loses $e^-$), usually a metal
- O.N. rules: free element = 0; monoatomic ion = its charge; Group1=+1, Group2=+2, Group13=+3; H=+1 (except hydrides -1); O=-2 (except peroxides -1, $OF_2$ +2)
- Electrolytic cell: non-spontaneous, uses current, anode(+)/cathode(−), electrical→chemical energy
- Galvanic/voltaic cell: spontaneous, produces current, anode(−)/cathode(+), chemical→electrical energy

- Anode = oxidation (in both cell types); Cathode = reduction (in both cell types)
- Downs cell → Na metal from fused NaCl; Nelson's cell → NaOH from brine
- Rusting: Fe oxidized at anodic region, O2 reduced at cathodic region (needs H+), rust = Fe2O3·nH2O
- Corrosion prevention: remove stains, paint/grease, alloying (stainless steel), metallic coating (galvanizing, tin, electroplating)
- Galvanizing protects even when scratched (Zn sacrificial); tin coating fails once broken (Fe becomes anode, rusts faster)

## Exam Tips

- Master the O.N. assignment rules cold -- this unit's numericals depend entirely on applying them correctly and quickly
- Always double-check which half-reaction is oxidation (electron loss, O.N. increases) vs reduction before labeling anode/cathode
- Memorize the anode/cathode charge reversal between electrolytic cells (anode +) and galvanic cells (anode −) -- a very common exam trap
- Draw and label the Daniel cell and Downs/Nelson's cells from memory -- diagram-based questions are common in this unit
- Learn the exact half-cell reactions for Downs cell, Nelson's cell, rusting, and at least one electroplating example (silver or chromium)
- Understand WHY galvanizing outperforms tin coating -- this reasoning-based question appears frequently
- Practice identifying oxidizing vs reducing agents by determining which species is reduced (oxidizing agent) and which is oxidized (reducing agent) in a given equation

# Unit 8: Chemical Reactivity

Chemical reactivity describes how readily elements gain, lose, or share electrons to form compounds. Metals are electropositive elements (except hydrogen) that form cations by losing electrons, while non-metals are electronegative elements that form anions by gaining electrons. This unit examines the reactivity trends of metals -- particularly alkali and alkaline earth metals -- and non-metals, especially the halogens.

Topics covered include the classification of metals by reactivity, electropositive character and its relationship to ionization energy, a detailed comparison of alkali metals and alkaline earth metals, the inertness of noble metals like silver, gold, and platinum, the physical and chemical properties of non-metals, and the comparative reactivity and important reactions of the halogens.

## Learning Objectives

- Relate the terms cation and anion to metals and non-metals respectively
- Explain why alkali metals are not found in the free state in nature
- Explain the differences in ionization energies of alkali and alkaline earth metals
- Describe the position of sodium in the periodic table along with its properties and uses
- Describe the position of calcium and magnesium in the periodic table along with their properties and uses
- Differentiate between soft and hard metals using iron and sodium as examples
- Describe the inertness of noble metals and identify the commercial value of silver, gold, and platinum
- Compile important reactions of the halogens
- Name elements that exist in nature in uncombined (free) form

## Key Concepts

### 8.1 Metals

Metals are elements (except hydrogen) that are electropositive and form cations by losing electrons. They are classified by reactivity into three groups: very reactive (potassium, sodium, calcium, magnesium, aluminium), moderately reactive (zinc, iron, tin, lead), and least reactive or noble (copper, mercury, silver, gold). Important physical properties of metals include: almost all are solids except mercury; they have high melting and boiling points (except alkali metals); they possess metallic lustre and can be polished; they are malleable and ductile and give a ringing tone when struck; they are good conductors of heat and electricity; they have high densities; and they are hard except sodium and potassium.

Important chemical properties of metals: they easily lose electrons to form positive ions; they react with oxygen to form basic oxides; they usually form ionic compounds with non-metals; and they exhibit metallic bonding. Notable facts: aluminium is the most abundant metal, platinum the most precious, iron the most useable, cesium the most reactive, lithium the lightest (0.53 $g/cm^3$), osmium the heaviest (22.5 $g/cm^3$), lead the poorest heat conductor, and silver and gold the best conductors and the most ductile/malleable metals.

### 8.1.1 Electropositive Character

Electropositivity (metallic character) is a metal's tendency to lose its valence electrons -- the more easily a metal loses electrons, the more electropositive it is. The number of electrons lost is the metal's valency; for example, sodium loses 1 electron ($Na \rightarrow Na^+ + e^-$, valency 1) while zinc loses 2 electrons ($Zn \rightarrow Zn^{2+} + 2e^-$, valency 2). Electropositivity increases down a group (as atomic size increases, e.g. Li < Na < K) and decreases across a period from left to right (as atomic size decreases due to increasing nuclear charge).

Electropositivity depends on ionization energy, which in turn depends on atomic size and nuclear charge: small atoms with high nuclear charge have high ionization energy and are therefore less electropositive. Alkali metals have the largest atomic size and lowest ionization energy in their periods, giving them the highest metallic character. For example, sodium (atomic size 186 pm, IE 496 kJ/mol) is more electropositive than magnesium (atomic size 160 pm, IE1 738 kJ/mol) -- and magnesium's second ionization energy (1450 kJ/mol) is far higher than its first, because removing an electron from the already-positive $Mg^+$ ion against strong nuclear attraction is much harder.

### 8.1.2 Comparison of Alkali and Alkaline Earth Metals

Group 1 elements are called alkali metals and Group 2 elements are called alkaline earth metals. Alkali metals have an $ns^1$ valence configuration, so their single valence electron is easily lost, making them extremely reactive and always found in nature combined as +1 cations. Alkaline earth metal atoms are smaller with two valence electrons ($ns^2$) and more nuclear charge, making them reactive but less so than alkali metals.

Comparing reactivity: alkali metals react vigorously with water at room temperature to give a strong alkaline solution and hydrogen gas ($2Na + 2H_2O \rightarrow 2NaOH + H_2$), while alkaline earth metals react less vigorously and need heating to produce weaker bases ($Mg + H_2O \rightarrow MgO + H_2$, then $MgO + H_2O \rightarrow Mg(OH)_2$). Alkali metals tarnish immediately in air forming strongly alkaline oxides ($4Na + O_2 \rightarrow 2Na_2O$), while alkaline earth metals need heating to react with oxygen ($2Mg + O_2 \rightarrow 2MgO$). Alkali metals react violently with halogens at room temperature ($2Na + Cl_2 \rightarrow 2NaCl$), while alkaline earth metals react more slowly ($Ca + Cl_2 \rightarrow CaCl_2$). Alkali metals do not form nitrides or carbides directly, while alkaline earth metals form stable nitrides and carbides on heating ($3Mg + N_2 \rightarrow Mg_3N_2$; $Ca + 2C \rightarrow CaC_2$).

Sodium is used in sodium-potassium alloy as a nuclear reactor coolant, in sodium vapour lamps for yellow light, and as a reducing agent in extracting metals like titanium. Magnesium is used in flashbulbs and fireworks, in light alloys, in the thermite process, and as a sacrificial anode for corrosion prevention. Calcium is used to remove sulphur from petroleum products and as a reducing agent to produce chromium, uranium, and zirconium.

### Inertness of Noble Metals

Transition metals (d-block elements) exhibit variable oxidation states; among Group 11, copper behaves like other active metals, but silver and gold are relatively inactive because they do not lose electrons easily. Silver is a white, lustrous, highly ductile and malleable metal, an excellent conductor of heat and electricity; a thin oxide/sulphide layer makes it relatively unreactive, though it tarnishes with sulphur-containing compounds like $H_2S$. Being soft, it is alloyed with copper for coins, silverware, and ornaments, and used in photographic films, dental preparations, and mirrors.

Gold is a soft, yellow, extremely malleable and ductile metal (one gram can be drawn into a 1.5 km wire) that is essentially inert -- unaffected by the atmosphere or by any single mineral acid or base. Its purity is measured in carats out of 24 (24 carat is pure gold); it is typically alloyed with copper, silver, palladium, nickel, or zinc for jewelry. Platinum is used in jewelry for its color, strength, and tarnish resistance, and (alloyed with palladium and rhodium) as a catalytic converter that converts toxic vehicle emissions ($CO$, $NO_2$) into less harmful $CO_2$, $N_2$, and water vapour; it is also used in hard disk coatings, fibre optics, and LCD glass manufacturing.

### 8.2 Non-Metals

Non-metals form negative ions (anions) by gaining electrons, making them electronegative, and they form acidic oxides. Valency depends on the number of electrons accepted -- chlorine accepts 1 electron (valency 1), oxygen accepts 2 (valency 2). Non-metallic character depends on electron affinity and electronegativity: small, high-nuclear-charge atoms are strongly electronegative, so non-metallic character decreases down a group and increases across a period up to the halogens; fluorine is the most non-metallic element. Non-metals occupy Group 14 (carbon), Group 15 (nitrogen, phosphorus), Group 16 (oxygen, sulphur, selenium), and Group 17 (the halogens).

Physical properties of non-metals: solid non-metals are brittle; they are bad conductors of heat and electricity (except graphite); they are dull, not shiny (except iodine, which is lustrous); they are generally soft (except diamond); they have low melting/boiling points (except silicon, graphite, and diamond); and they have low densities. Chemical properties: their valence shells are electron-deficient, so they readily accept electrons; they form ionic compounds with metals and covalent compounds with other non-metals; they usually do not react with water; and they do not react with dilute acids since they are themselves electron acceptors. Electronegativity decreases in the order F > O > Cl > N > Br > S > C > I > P.

### 8.2.1 & 8.2.2 Halogen Reactivity and Reactions

Group 17 (halogens: fluorine, chlorine, bromine, iodine, astatine) have valence configuration $ns^2np^5$, needing only one more electron to complete their octet -- so they readily accept an electron from metals (forming ionic bonds) or share an electron with non-metals (forming covalent bonds). Fluorine and chlorine are diatomic gases at room temperature; as atomic size (and intermolecular forces) increases down the group, bromine is a liquid and iodine a solid.

All halogens are oxidizing agents, with strength decreasing down the group: fluorine is the strongest oxidizer, iodine the weakest. A more reactive halogen displaces a less reactive one from its salt solution -- e.g. $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$ (solution turns reddish brown) and $Br_2 + 2KI \rightarrow 2KBr + I_2$. All halogens combine with hydrogen to form hydrogen halides, but reactivity decreases down the group: fluorine reacts even in the dark and cold, chlorine needs sunlight, and bromine/iodine need heating (with iodine also needing a catalyst). Similarly, fluorine decomposes water in cold and dark conditions, chlorine needs sunlight, bromine reacts only under special conditions, and iodine does not react with water at all.

With methane, fluorine reacts violently even in the dark, while chlorine only reacts (violently) in bright sunlight, or slowly in diffused light giving a series of substitution products ($CH_3Cl$, $CH_2Cl_2$, $CHCl_3$, $CCl_4$). Chlorine reacts with cold dilute NaOH to give sodium hypochlorite and sodium chloride ($2NaOH + Cl_2 \rightarrow NaCl + NaOCl + H_2O$), and with hot concentrated NaOH to give sodium chloride and sodium chlorate ($6NaOH + 3Cl_2 \rightarrow 5NaCl + NaClO_3 + 3H_2O$).

### 8.2.3 Significance of Non-metals

Non-metals, though fewer than metals, are vital to life: oxygen is the most abundant element in the earth's crust (47%) and oceans (86%); about 96% of human body mass is made up of just four non-metals -- oxygen (65%), carbon (18%), hydrogen (10%), and nitrogen (3%); $O_2$ and $CO_2$ are essential for respiration in animals and plants respectively; carbohydrates, proteins, fats, and water are all built from non-metals; and atmospheric nitrogen (78%) prevents uncontrolled combustion. Fossil fuels (coal, petroleum, gas) are composed of carbon and hydrogen, and everyday materials like natural/synthetic fibres, plastics, and agrochemicals are also built from non-metallic elements.

## Important Definitions

### Define metals in terms of ion formation.

Elements (except hydrogen) that are electropositive and form positive ions (cations) by losing electrons.

### Define non-metals in terms of ion formation.

Elements that are electronegative and form negative ions (anions) by gaining electrons.

### Define electropositivity (metallic character).

The tendency of a metal atom to lose its valence electrons; the more easily electrons are lost, the more electropositive the metal.

### Define valency of a metal.

The number of electrons an atom of the metal loses to form a positive ion.

### Define alkali metals.

The Group 1 elements, having an ns1 valence configuration, which are extremely reactive and always found combined in nature as +1 cations.

### Define alkaline earth metals.

The Group 2 elements, having an ns2 valence configuration, which are reactive but less so than alkali metals.

### Define noble metals.

Metals such as copper, mercury, silver, and gold that are the least reactive metals, resisting oxidation and corrosion.

### Define carat (in the context of gold purity).

A unit indicating the number of parts by weight of pure gold present in 24 parts of an alloy; 24 carat gold is pure gold.

### Define non-metallic character (electronegativity-related).

The tendency of an atom to attract and gain electrons, depending on its electron affinity and electronegativity.

### Define halogens.
The Group 17 elements (fluorine, chlorine, bromine, iodine, astatine), which have an ns2np5 valence configuration and readily gain one electron to complete their octet.

### Define oxidizing agent (in the context of halogens).
A species that oxidizes another substance by removing electrons from it; among halogens, fluorine is the strongest oxidizing agent and iodine the weakest.

### Define malleability.
The property of a metal that allows it to be hammered or rolled into thin sheets without breaking.

### Define ductility.
The property of a metal that allows it to be drawn into thin wires without breaking.

### Define brittleness.
The tendency of a solid (typically non-metals) to break or shatter easily rather than deform.

### Define transition metals.
Elements in which the d-orbitals are in the process of being filled, exhibiting variable oxidation states; they occupy the middle of periods 4, 5, and 6 of the periodic table.

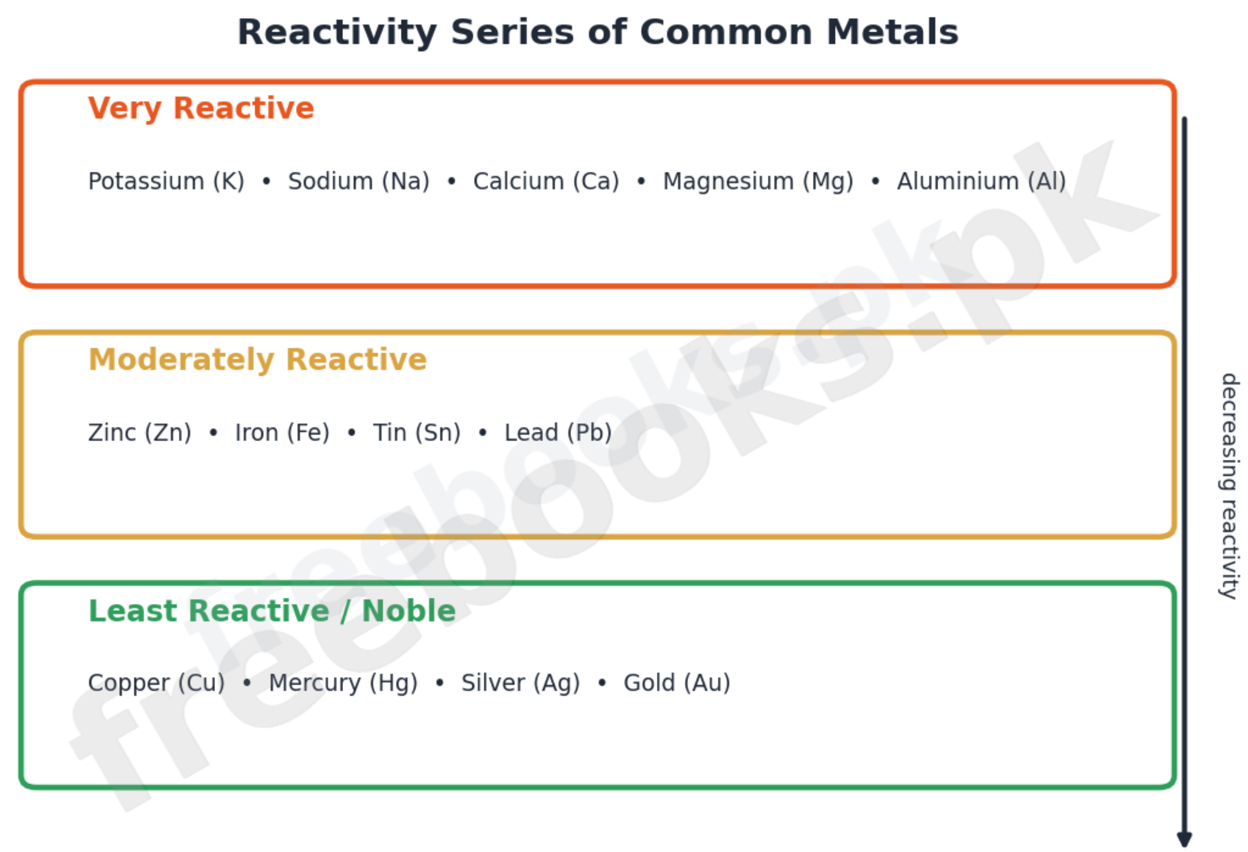
## Key Formulas

| Topic | Formula |
|---|---|
| Sodium losing an electron | $Na(s) \rightarrow Na^{+}(g) + 1e^{-}$ (valency 1) |
| Zinc losing electrons | $Zn(s) \rightarrow Zn^{2+}(g) + 2e^{-}$ (valency 2) |
| Chlorine gaining an electron | $Cl + 1e^{-} \rightarrow Cl^{-}$ (valency 1) |
| Oxygen gaining electrons | $O + 2e^{-} \rightarrow O^{2-}$ (valency 2) |
| Alkali metal + water | $2Na + 2H_2O \rightarrow 2NaOH + H_2$ |
| Alkaline earth metal + water | $Mg + H_2O \rightarrow MgO + H_2$ (then $MgO + H_2O \rightarrow Mg(OH)_2$) |

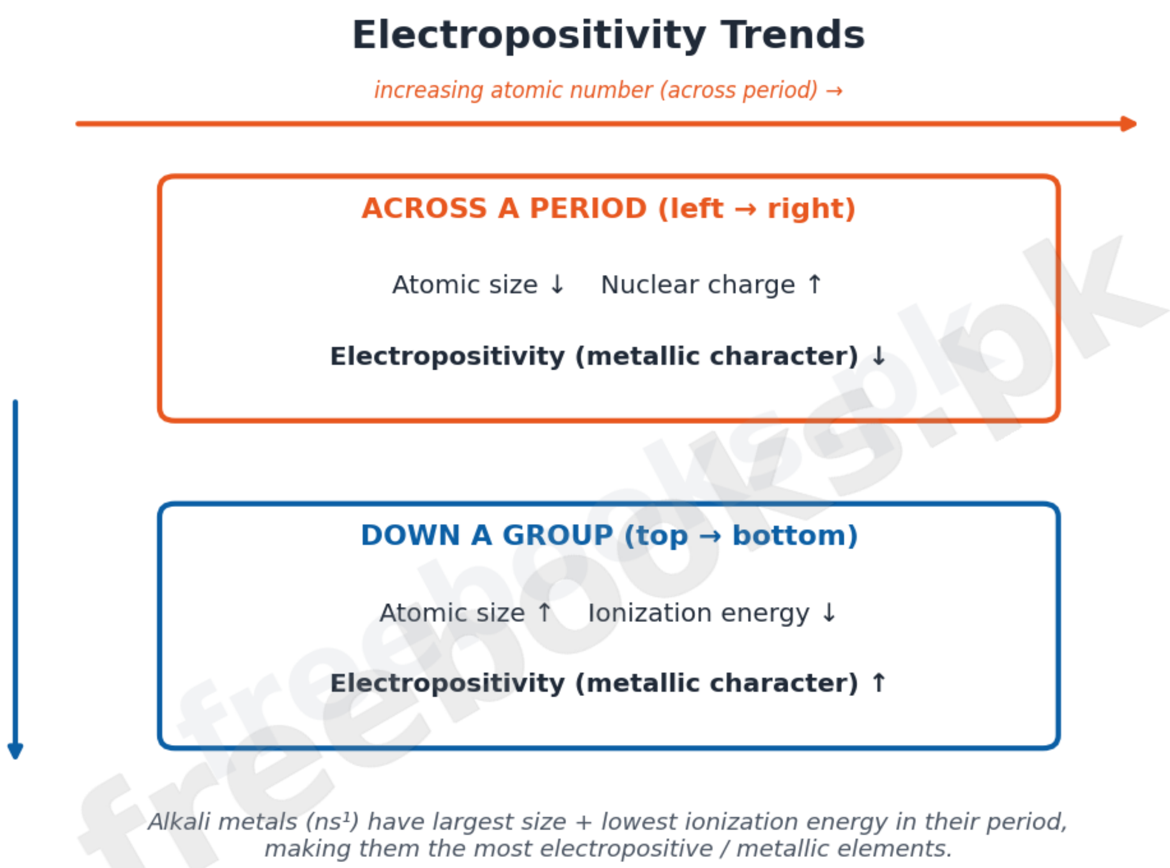
| | |
|---|---|
| Halogen displacement | $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$ |

## Diagrams

**Reactivity Series of Common Metals:** Classification of common metals into very reactive, moderately reactive, and least reactive (noble) categories.




**Electropositivity Trends:** Summary of how electropositivity (metallic character) changes across a period and down a group in the periodic table.



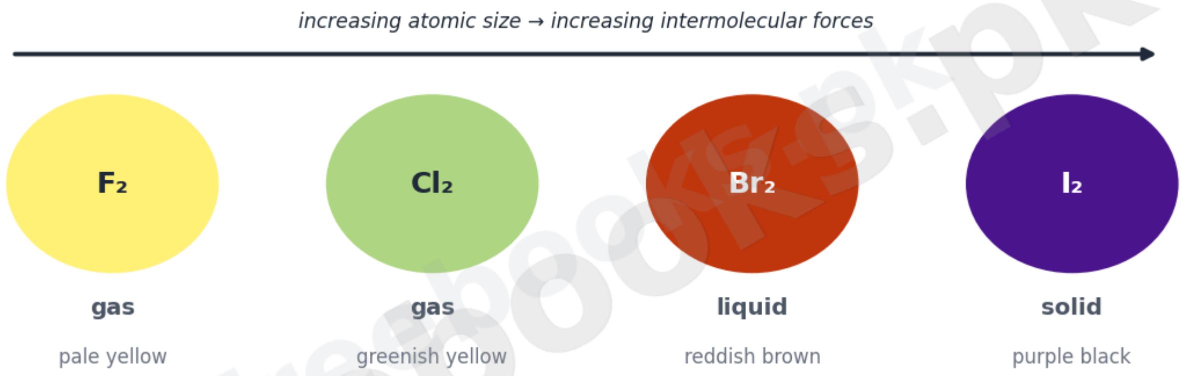

**Alkali Metals vs Alkaline Earth Metals: Reactivity:** Side-by-side comparison of alkali and alkaline earth metal reactions with water, oxygen, and halogens.

## Alkali Metals vs Alkaline Earth Metals: Reactivity

| | ALKALI METALS | ALKALINE EARTH METALS |
|---|---|---|
| Valence config | $ns^1$ (1 electron to lose) | $ns^2$ (2 electrons to lose) |
| Reaction with water | Vigorous at room temp: $2Na+2H_2O \rightarrow 2NaOH+H_2$ | Slower, needs heat: $Mg+H_2O \rightarrow MgO+H_2$ |
| Reaction with $O_2$ | Immediate tarnishing: $4Na+O_2 \rightarrow 2Na_2O$ | Needs heating: $2Mg+O_2 \rightarrow 2MgO$ |
| Reaction with halogens | Violent at room temp: $2Na+Cl_2 \rightarrow 2NaCl$ | Slower reaction: $Ca+Cl_2 \rightarrow CaCl_2$ |

**Halogen Reactivity and Physical State Trend:** Comparison of the four common halogens by physical state, colour, and oxidizing/hydrogen reactivity trend.

## Halogen Reactivity and Physical State Trend




**Oxidizing power: $F_2 > Cl_2 > Br_2 > I_2$ (fluorine strongest, iodine weakest)**

Reactivity with $H_2$: $F_2$ (dark & cold) > $Cl_2$ (sunlight) > $Br_2$ (heat) > $I_2$ (heat + catalyst)

## Short Questions & Answers

### Why is sodium metal more reactive than magnesium metal?

Because sodium has a larger atomic size and a much lower ionization energy than magnesium, so it loses its single valence electron far more easily, making it more electropositive and reactive.

### Why does the second ionization energy of magnesium far exceed its first?

Because after removing the first electron, the resulting $Mg^+$ ion has a stronger effective nuclear attraction on its remaining electrons (and a smaller size), making it much harder to remove the second electron.

### Why are alkali metals always found in combined form in nature rather than as free elements?

Because their single valence electron ($ns^1$ configuration) is lost extremely easily, making them highly reactive; they readily combine with other elements to form stable compounds rather than existing as free, unreacted metal.

### Why is gold considered a highly inert metal?

Because gold's electron configuration makes it very reluctant to lose electrons, so it resists reaction with atmospheric oxygen, moisture, and even individual mineral acids or bases under normal conditions.

### Why is pure gold rarely used to make jewelry?

Because pure gold is too soft and easily damaged for practical, long-lasting jewelry, so it is alloyed with metals like copper, silver, palladium, nickel, or zinc to increase its hardness and durability.

### Why do non-metals generally not react with dilute acids?

Because reacting with a dilute acid typically requires donating electrons to displace hydrogen, but non-metals are themselves electron acceptors (electronegative), so they lack the tendency to give up electrons in this way.

### Why does fluorine react with hydrogen even in the dark and cold, while iodine needs heat and a catalyst?

Because chemical affinity for hydrogen decreases down the halogen group as atomic size increases and bond energies change, making fluorine far more reactive under mild conditions and iodine reactive only under forcing conditions.

### Why is bromine a liquid and iodine a solid at room temperature, while fluorine and chlorine are gases?

Because intermolecular (van der Waals) forces of attraction increase down the halogen group as atomic and molecular size increases, making the heavier halogens more strongly attracted to each other and less volatile.

### Why does the oxidizing power of halogens decrease from fluorine to iodine?

Because oxidizing power depends on the ability to attract and gain an electron, which is strongest for fluorine (smallest size, highest electronegativity) and progressively weaker for the larger, less electronegative halogens down the group.

### Why is oxygen considered one of the most important non-metals for life?

Because oxygen is the most abundant element in the earth's crust and oceans, makes up a large fraction of human body mass, and is essential for respiration in animals -- without it, life as we know it would not be possible.

## Long Questions & Answers

### Compare and contrast the properties and reactivity of alkali metals and alkaline earth metals.

#### How do the valence electron configurations of alkali and alkaline earth metals differ, and what effect does this have?

Alkali metals (Group 1) have an ns1 configuration with a single, easily-lost valence electron, making them extremely reactive and always found combined in nature. Alkaline earth metals (Group 2) have an ns2 configuration with two valence electrons and greater nuclear charge, making them reactive but noticeably less so than alkali metals.

#### How do the two groups differ in their reaction with water?

Alkali metals react vigorously with water even at room temperature, producing a strongly alkaline solution and hydrogen gas (e.g. $2Na + 2H_2O \rightarrow 2NaOH + H_2$). Alkaline earth metals react much less vigorously, typically requiring heat, and produce weaker bases (e.g. $Mg + H_2O \rightarrow MgO + H_2$, followed by $MgO + H_2O \rightarrow Mg(OH)_2$).

### How do the two groups differ in their reaction with oxygen and halogens?

Alkali metals tarnish almost immediately in air, forming strongly alkaline oxides ($4Na + O_2 \rightarrow 2Na_2O$), and react violently with halogens even at room temperature ($2Na + Cl_2 \rightarrow 2NaCl$). Alkaline earth metals need heating to react with oxygen ($2Mg + O_2 \rightarrow 2MgO$) and react more slowly with halogens ($Ca + Cl_2 \rightarrow CaCl_2$).

### How do the two groups differ in their reaction with nitrogen and carbon?

Alkali metals do not form nitrides or carbides directly under normal conditions. Alkaline earth metals, in contrast, form stable nitrides and carbides when heated with nitrogen or carbon respectively (e.g. $3Mg + N_2 \rightarrow Mg_3N_2$ and $Ca + 2C \rightarrow CaC_2$), reflecting their comparatively different bonding behaviour despite both being highly reactive metal groups.

## Discuss the inert character of silver and gold, and explain why they are valued as noble metals.

### Why are silver and gold classified as noble (inert) metals?

Silver and gold, along with copper, belong to Group 11 of the transition metals, but unlike copper, silver and gold do not lose electrons easily. This reluctance to react makes them chemically inactive compared to most other metals, which is why they are described as noble or inert metals.

### What specifically makes silver relatively unreactive, and where does it still react?

Silver forms a thin protective layer of oxide or sulphide on its surface, which makes it largely unaffected by normal atmospheric conditions. However, it does tarnish in the presence of sulphur-containing compounds such as hydrogen sulphide ($H_2S$), showing that its inertness is not absolute.

**How inert is gold compared to silver, and what practical consequence does this have?**

Gold is even more inert than silver: it is not affected by the atmosphere and does not react with any single mineral acid or base under normal conditions. Because of this exceptional resistance to corrosion and tarnishing, gold retains its lustre and appearance indefinitely, making it highly prized for ornamental use and as a stable store of value.

**Why must both silver and gold typically be alloyed before practical use?**

Both metals are naturally very soft in their pure form, making them impractical for jewelry, coins, or tools on their own. Silver is commonly alloyed with copper, while gold is alloyed with copper, silver, or other metals (with purity expressed in carats out of 24), giving both metals the added hardness and durability needed for everyday and ornamental use while retaining their prized inertness and appearance.

## Multiple Choice Questions (MCQs)

### Metals can form ions carrying charges: (A) uni-positive (B) di-positive (C) tri-positive (D) all of these

Correct answer: (D) all of these. Different metals lose different numbers of electrons depending on their valency, so metals collectively can form uni-positive, di-positive, or tri-positive ions.

### Which one of the following metals burns with a brick red flame? (A) sodium (B) magnesium (C) iron (D) calcium

Correct answer: (D) calcium. Calcium characteristically produces a brick red flame when burned, as noted in its comparison with sodium (golden yellow) and magnesium (brilliant white).

### Sodium is an extremely reactive metal, but it does not react directly with: (A) hydrogen (B) nitrogen (C) sulphur (D) phosphorus

Correct answer: (B) nitrogen. Alkali metals like sodium do not form nitrides directly, unlike alkaline earth metals, which do form stable nitrides on heating with nitrogen.

**Which one of the following is the lightest metal? (A) calcium (B) magnesium (C) lithium (D) sodium**

Correct answer: (C) lithium. Lithium has the lowest density (0.53 g/cm3) among common metals, making it the lightest metal listed.

**Pure alkali metals can be cut simply with a knife but iron cannot, because alkali metals have: (A) strong metallic bonding (B) weak metallic bonding (C) non-metallic bonding (D) moderate metallic bonding**

Correct answer: (B) weak metallic bonding. Alkali metals have relatively weak metallic bonding due to their large atomic size and single valence electron, making them soft enough to cut with a knife, unlike strongly-bonded, hard metals like iron.

**Which one of the following is less malleable? (A) sodium (B) iron (C) gold (D) silver**

Correct answer: (B) iron. Iron is comparatively less malleable than the very soft alkali metal sodium and the exceptionally malleable noble metals gold and silver.

**Metals lose their electrons easily because: (A) they are electronegative (B) they have electron affinity (C) they are electropositive (D) they are good conductors of heat**

Correct answer: (C) they are electropositive. Metals are described as electropositive precisely because of their tendency to lose electrons and form positive ions, which is the underlying reason for this behaviour.

**Which one of the following is brittle? (A) sodium (B) aluminium (C) selenium (D) magnesium**

Correct answer: (C) selenium. Selenium, a non-metal, is brittle like most solid non-metals, unlike the metals sodium, aluminium, and magnesium, which are malleable.

**Which one of the following non-metals is lustrous? (A) sulphur (B) phosphorus (C) iodine (D) carbon**

Correct answer: (C) iodine. Iodine is unusual among non-metals in having a lustrous, shiny appearance similar to metals, despite being a non-metal.

**Non-metals are generally soft, but which one of the following is extremely hard? (A) graphite (B) phosphorus (C) iodine (D) diamond**

Correct answer: (D) diamond. Diamond, an allotrope of carbon, is an exception among non-metals, being one of the hardest known natural materials.

## Quick Revision Summary

- Metals = electropositive, form cations (lose $e^-$); Non-metals = electronegative, form anions (gain $e^-$)
- Metal reactivity groups: very reactive (K,Na,Ca,Mg,Al) > moderately reactive (Zn,Fe,Sn,Pb) > noble (Cu,Hg,Ag,Au)
- Electropositivity: increases down a group, decreases across a period (opposite trend to non-metallic character)
- Electropositivity ∝ 1/ionization energy; alkali metals have lowest IE, highest metallic character
- Alkali metals ($ns^1$): very reactive, always combined in nature; Alkaline earth metals ($ns^2$): reactive but less so
- Silver and gold are 'noble' -- inert due to reluctance to lose electrons; gold most inert, doesn't react with acids/bases
- Gold purity in carats (24 = pure); both Ag and Au alloyed for hardness in jewelry/coins
- Non-metals: Group14 (C), Group15 (N,P), Group16 (O,S,Se), Group17 (halogens); non-metallic character ∝ electronegativity
- Halogens ($ns^2np^5$): oxidizing power $F_2>Cl_2>Br_2>I_2$; reactivity with $H_2$ and $H_2O$ both decrease down the group
- Halogen physical state trend: $F_2$,$Cl_2$ gases; $Br_2$ liquid; $I_2$ solid (increasing intermolecular forces down group)

## Exam Tips

- Memorize the three metal reactivity categories (very reactive / moderately reactive / noble) with example metals in each
- Be ready to explain the electropositivity-ionization energy relationship with the Na vs Mg comparison as a worked example
- Learn all six alkali vs alkaline earth metal reaction comparisons (water, O2, H2, halogens, N2, C) -- these are exam favourites
- Know specific uses of sodium, magnesium, and calcium -- short-answer questions often ask for these directly

- Understand WHY gold and silver are inert (electron-loss reluctance) rather than just memorizing that they are
- Memorize the halogen reactivity trends (oxidizing power, reaction with $H_2$/$H_2O$/methane) in the correct decreasing order
- Practice distinguishing metal vs non-metal properties side by side, since comparison questions are common in this unit