SUMMARY OF STRANDS AND SUB STRANDS

1.0 INORGANIC CHEMISTRY

1.1. Introduction Chemistry

1.2. The Atom

1.3. The periodic Table

1.4. Chemical Bonding

1.5. Periodicity

2.0 PHYSICAL CHEMISTRY

2.1. Acids and Bases

2.2. Introduction to Salts

3.0 ORGANIC CHEMISTRY

STRAND 1.0: INORGANIC CHEMISTRY

Sub-Strand 1.1: Introduction to Chemistry

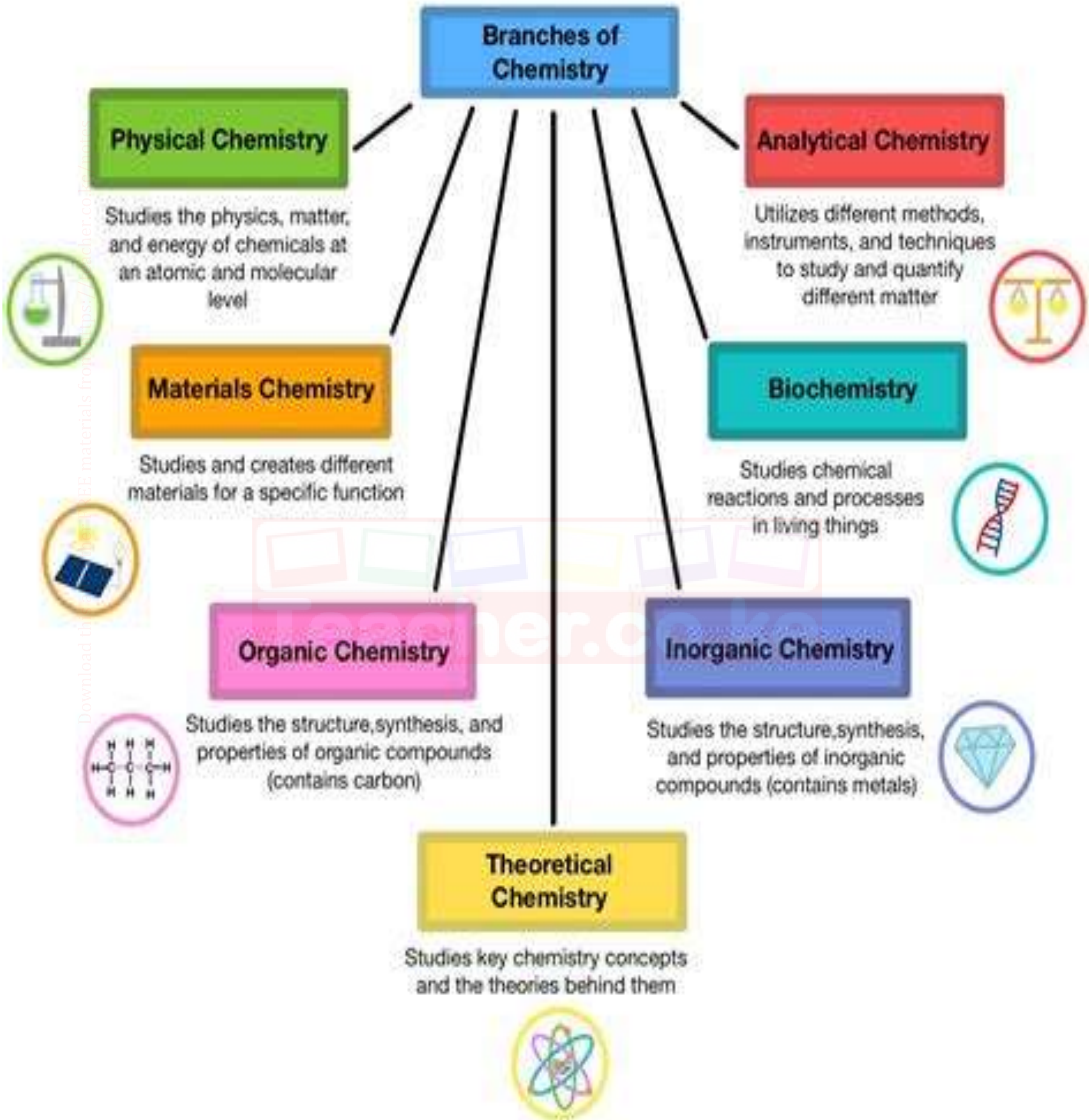
(a) Meaning of Chemistry as a Field of Science

Chemistry is a **branch of science** that studies **matter** and its **properties**, as well as how matter **changes** and **interacts** with energy. It focuses on the fundamental building blocks of matter, which are **atoms**, and how these atoms combine to form **molecules** and **compounds**. Chemistry seeks to understand the composition, structure, properties, and reactions of these substances.

(b) Branches of Chemistry

Chemistry is a broad field with many specialized areas. Some of the main branches include:

- ✓ **Organic Chemistry:** The study of carbon-containing compounds (with some exceptions). This branch is crucial for understanding life processes and creating many synthetic materials.
- ✓ **Inorganic Chemistry:** The study of compounds that generally do not contain carbon. This includes metals, minerals, and many industrial chemicals..
- ✓ **Physical Chemistry:** This branch deals with the principles of physics that underlie chemical interactions and processes. It involves studying energy changes, reaction rates, and the structure of matter.
- ✓ **Analytical Chemistry:** This involves identifying and quantifying the components of substances. It plays a vital role in quality control, environmental monitoring, and medical diagnostics.
- ✓ **Biochemistry:** This is the study of chemical processes within living organisms. It explores the chemistry of proteins, carbohydrates, lipids, and nucleic acids.
- ✓ **Industrial Chemistry:** This branch focuses on the chemical processes involved in manufacturing products on a large scale, from pharmaceuticals to plastics.


Branches of Chemistry
Physical Chemistry
Studies the physics, matter, and energy of chemicals at an atomic and molecular level
Analytical Chemistry
Utilizes different methods, instruments, and techniques to study and quantify different matter
Materials Chemistry
Studies and creates different materials for a specific function
Biochemistry
Studies chemical reactions and processes in living things
Organic Chemistry
Studies the structure,synthesis, and properties of organic compounds (contains carbon)
Inorganic Chemistry
Studies the structure,synthesis, and properties of inorganic compounds (contains metals)
Theoretical Chemistry
Studies key chemistry concepts and the theories behind them

(c) Chemistry in Our Daily Lives

Chemistry plays a fundamental role in almost every aspect of our lives:

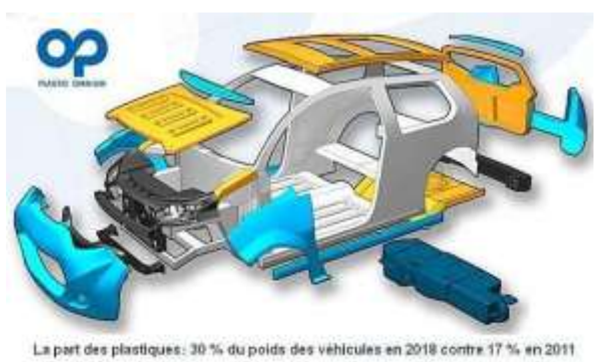
- ✓ **Food:** Chemistry helps us understand food composition, preservation, and digestion. Fertilizers and pesticides, developed through chemistry, increase food production.

- ✓ **Medicine:** The development of new drugs and pharmaceuticals is heavily reliant on chemistry. Understanding chemical reactions in the body helps treat diseases.

- ✓ **Clothing:** Many fabrics we wear are synthetic polymers developed through chemical processes. Dyes used to color clothes are also products of chemistry.

- **Household Products:** Soaps, detergents, cleaning agents, and cosmetics are all formulated using chemical principles.

- **Transportation:** Fuels like petrol and diesel are hydrocarbons, and the development of alternative fuels like biofuels involves chemistry. Materials used in vehicles, such as plastics and alloys, are also products of chemical research.



- **Agriculture:** Fertilizers, herbicides, and pesticides are chemically formulated to improve crop yields and protect plants from diseases and pests.

- **Energy:** Chemistry is central to energy production, from burning fossil fuels to developing renewable energy sources like solar panels and batteries.

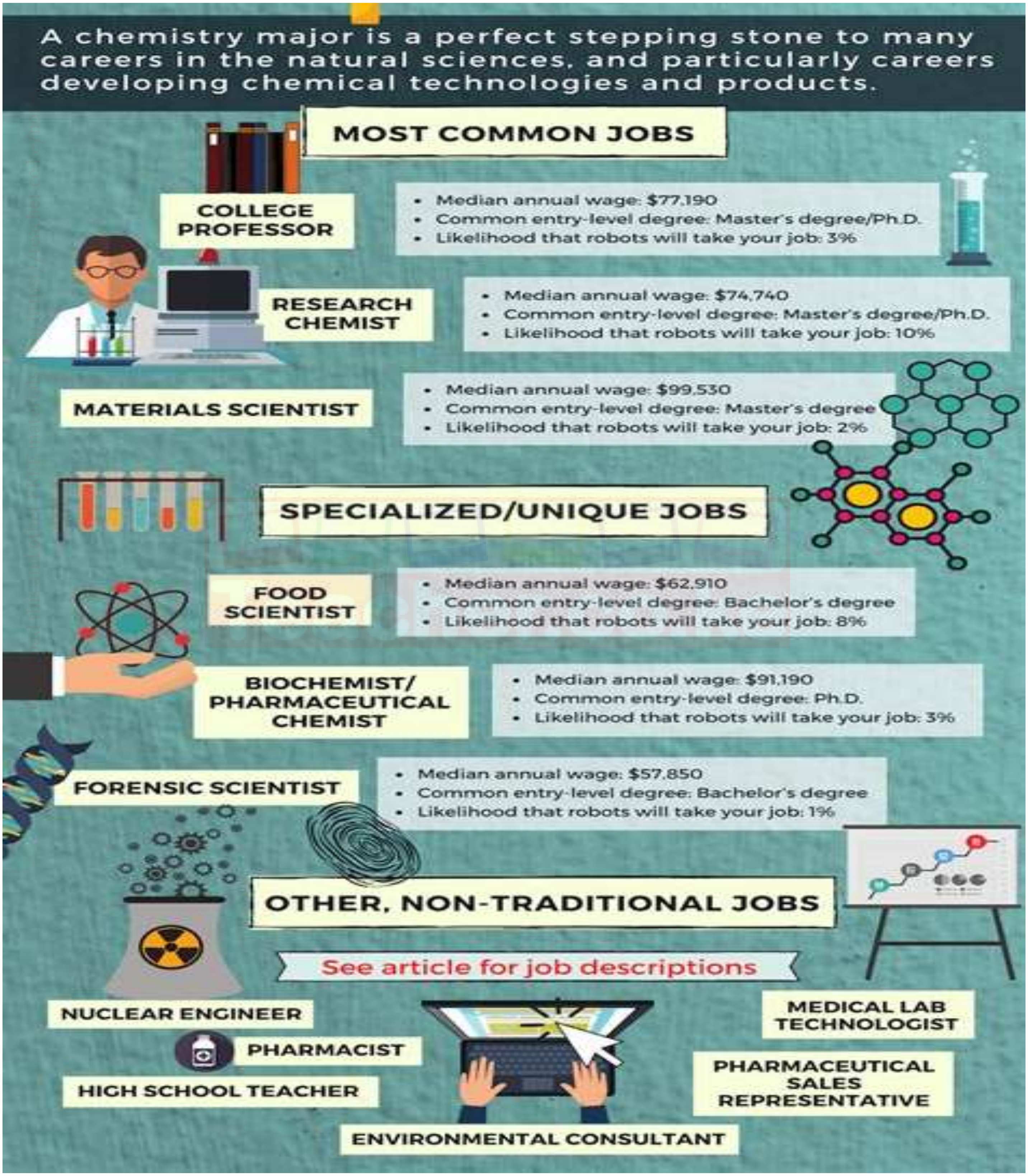
- **Environment:** Chemistry helps us understand and address environmental issues like pollution and climate change. It is used in developing solutions for waste management and water purification.

- Cooking and food preservation
- Cleaning products
- Medicines and healthcare
- Fuels and energy
- Plastic and textile industries
- Agriculture (fertilizers, pesticides)
- Cosmetics

Careers in Chemistry: Studying chemistry opens doors to a wide range of exciting career paths in Kenya and globally, including:

- **Chemical Engineer:** Designs and operates chemical plants.
- **Analytical Chemist:** Analyzes substances to determine their composition. (e.g., at the Kenya Revenue Authority (KRA) as seen in recent job postings)..
- **Biochemist:** Studies the chemistry of living organisms.
- **Forensic Scientist**
- **Pharmacist:** Dispenses and advises on medications.
- **Materials Scientist:** Develops and tests new materials.
- **Food Scientist:** Works on the chemistry of food production and safety.
- **Environmental Chemist:** Studies the chemical processes in the environment and works on pollution control.
- **Research Chemist:** Conducts experiments to discover new chemical knowledge.
- **Teacher/Lecturer:** Educates future generations about chemistry.
- **Lab Technician:** Assists scientists in conducting experiments and analyses.

Visual Aid:



A chemistry major is a perfect stepping stone to many careers in the natural sciences, and particularly careers developing chemical technologies and products.
MOST COMMON JOBS
COLLEGE PROFESSOR
• Median annual wage: $77,190
• Common entry-level degree: Master's degree/Ph.D.
• Likelihood that robots will take your job: 3%
RESEARCH CHEMIST
• Median annual wage: $74,740
• Common entry-level degree: Master's degree/Ph.D.
• Likelihood that robots will take your job: 10%
MATERIALS SCIENTIST
• Median annual wage: $99,530
• Common entry-level degree: Master's degree
• Likelihood that robots will take your job: 2%
SPECIALIZED/UNIQUE JOBS
FOOD SCIENTIST
• Median annual wage: $62,910
• Common entry-level degree: Bachelor's degree
• Likelihood that robots will take your job: 8%
BIOCHEMIST/ PHARMACEUTICAL CHEMIST
• Median annual wage: $91,190
• Common entry-level degree: Ph.D.
• Likelihood that robots will take your job: 3%
FORENSIC SCIENTIST
• Median annual wage: $57,850
• Common entry-level degree: Bachelor's degree
• Likelihood that robots will take your job: 1%
OTHER, NON-TRADITIONAL JOBS
See article for job descriptions
NUCLEAR ENGINEER
PHARMACIST
HIGH SCHOOL TEACHER
ENVIRONMENTAL CONSULTANT
MEDICAL LAB TECHNOLOGIST
PHARMACEUTICAL SALES REPRESENTATIVE

d) Effects of Drug and Substance Use in Day-to-Day Life

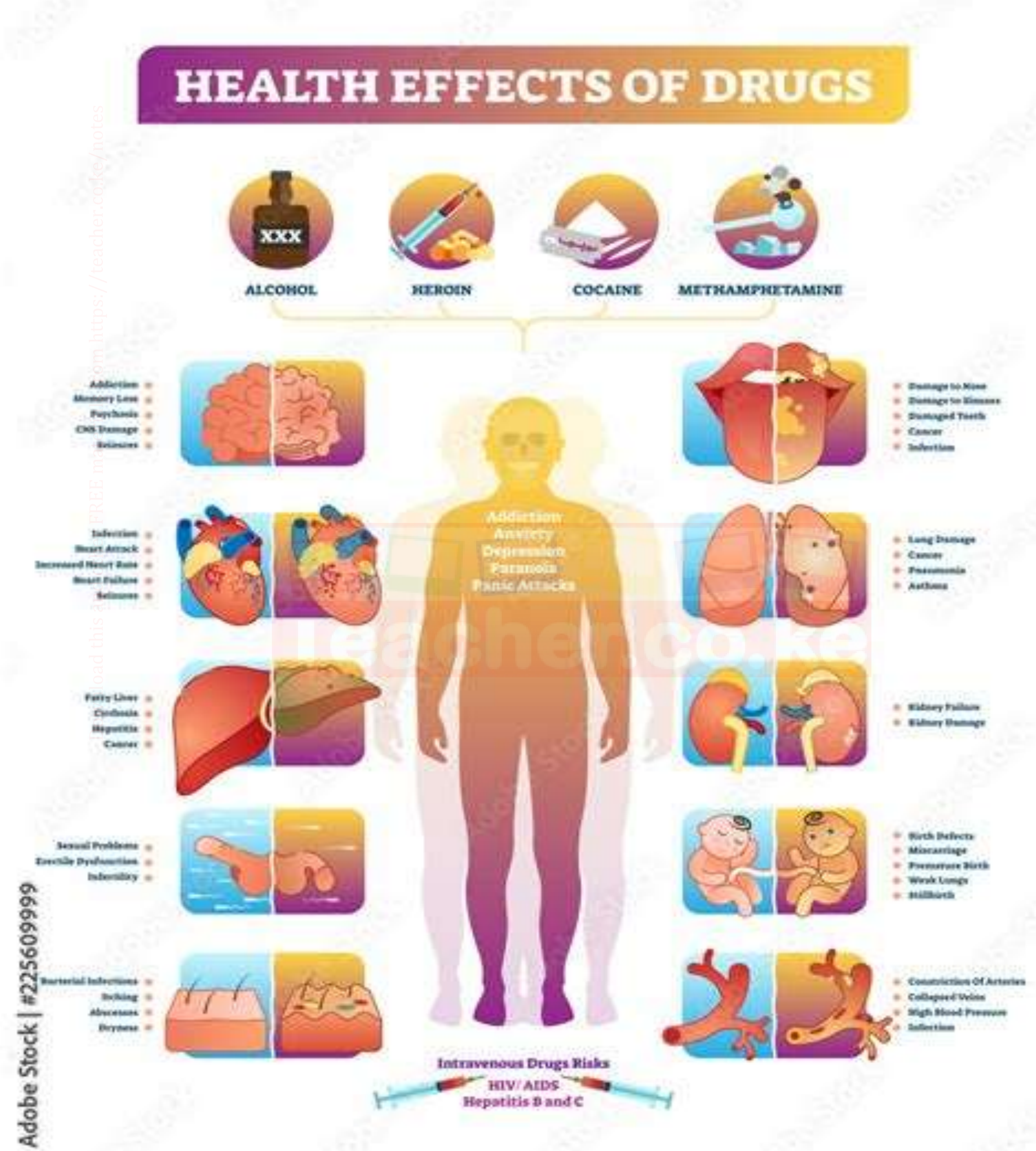
A **drug** is any substance that causes a change in an organism's physiology or psychology when consumed.

Drug and substance use can have severe negative effects on individuals and society:

- ✓ **Health Problems:** Substance abuse can lead to various physical and mental health issues, including addiction, organ damage, mental disorders like depression and anxiety, and even death.
- ✓ **Academic and Work Performance:** Drug and substance use can impair concentration, memory, and motivation, leading to poor performance in school or work, and potentially job loss..
- ✓ **Social Problems:** Substance abuse can strain relationships with family and friends, lead to social isolation, and increase the risk of criminal activity.
  Economic Burden: Drug and substance abuse can lead to significant financial problems for individuals and families due to the cost of drugs, treatment, and legal issues. It also places a burden on society through increased healthcare costs and reduced productivity.

- ✓ Mental health issues (e.g., anxiety, depression)
- ✓ Poor academic performance
- ✓ Addiction
- ✓ Risky behavior
- ✓ Diseases such as liver damage and heart disease

Visual Aid:


HEALTH EFFECTS OF DRUGS
XXX
ALCOHOL
HEROIN
COCAINE
METHAMPHETAMINE
Addiction
Memory Loss
Psychosis
CNS Damage
Seizures
Infection
Heart Attack
Increased Heart Rate
Heart Failure
Seizures
Fatty Liver
Cirrhosis
Hepatitis
Cancer
Sexual Problems
Erectile Dysfunction
Infertility
Bacterial Infections
Itching
Abscesses
Dryness
Addiction
Anxiety
Depression
Paranoia
Panic Attacks
Damage to Nose
Damage to Sinuses
Damaged Teeth
Cancer
Infection
Lung Damage
Cancer
Pneumonia
Asthma
Kidney Failure
Kidney Damage
Birth Defects
Miscarriage
Premature Birth
Weak Lungs
Stillbirth
Constriction Of Arteries
Collapsed Veins
High Blood Pressure
Infection
Intravenous Drugs Risks
HIV/ AIDS
Hepatitis B and C
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(e) Promoting Rights and Responsibilities to a Safe and Healthy Learning Environment

It is our right and responsibility to ensure a safe and healthy learning environment. This includes:

- ✓ **Staying informed:** Learning about the risks and effects of drug and substance use.
- ✓ **Making healthy choices:** Avoiding drug and substance use.
- ✓ **Seeking help:** Encouraging peers and seeking help if struggling with substance use or witnessing others who are.
- ✓ **Respecting oneself and others:** Promoting a culture of health and well-being.
- ✓ **Understanding consumer rights:** Being aware of regulations related to drug prescription and usage to ensure safety.

Promoting Safe and Healthy Learning

Learners should:

- ✓ Respect school rules and policies
- ✓ Avoid harmful substances
- ✓ Report unsafe conditions
- ✓ Support one another's health and well-being
- ✓ Promote gender equality in science careers

Project:

Develop posters or presentations to sensitize your peers and the community on the risks associated with drug and substance use and promote healthy lifestyles.

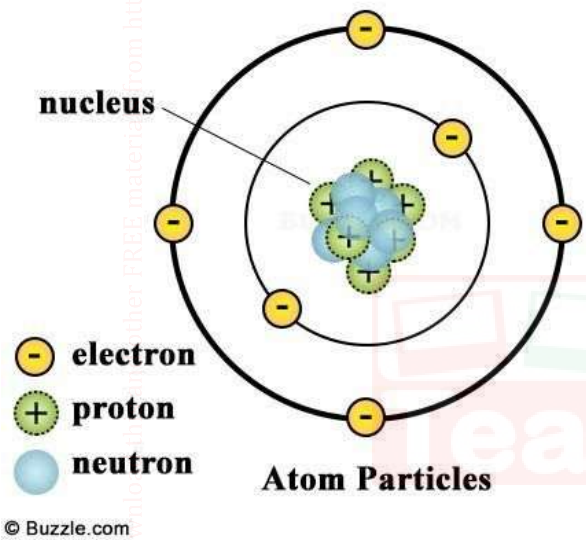
Suggestion for Project:

Sub-Strand 1.2: The Atom

(a) Structure of the Atom

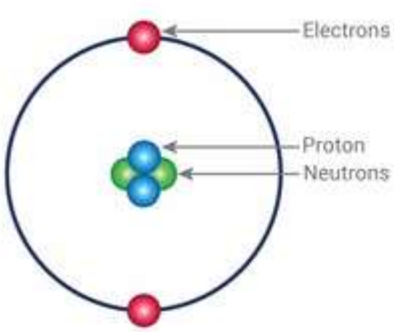
The atom is the basic unit of a chemical element. It consists of a dense central **nucleus** surrounded by a cloud of negatively charged **electrons**. The nucleus contains:

- **Protons:** Positively charged particles. The number of protons in an atom determines its **atomic number (Z)** and defines the element.
- **Neutrons:** Neutrally charged particles. The number of protons and neutrons together determines the **mass number (A)** of an atom.



(i) Dalton's Atomic Model (Brief Review)

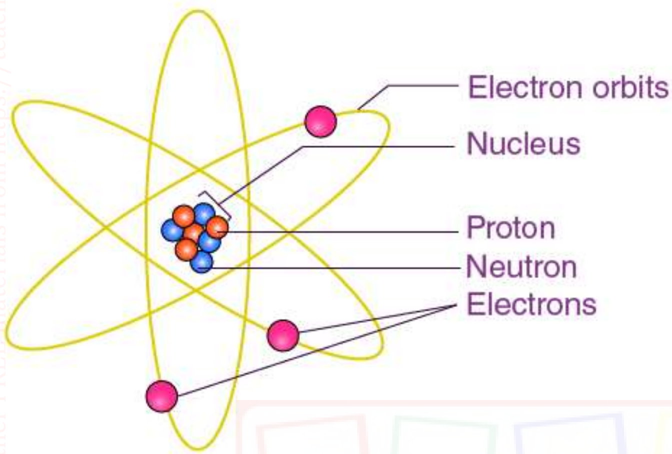
John Dalton proposed that all matter is made of tiny, indivisible particles called atoms. He suggested that atoms of the same element are identical and that chemical reactions involve the rearrangement of atoms.

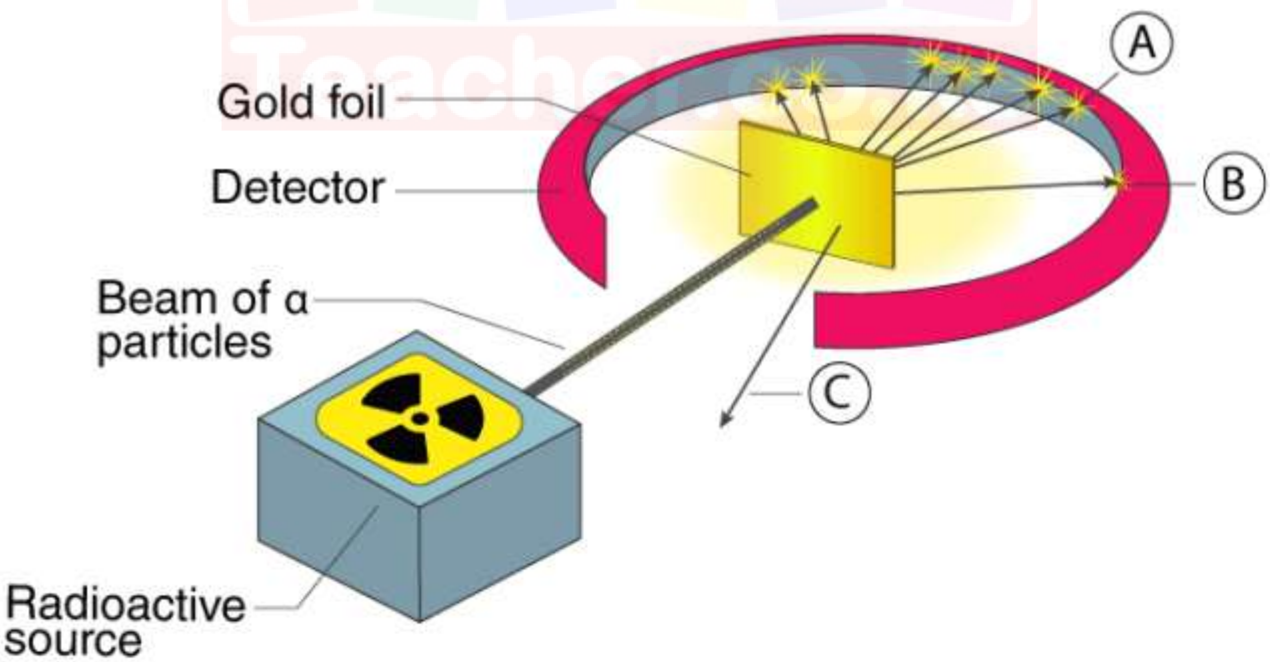


(ii) Rutherford's Atomic Model

Ernest Rutherford's gold foil experiment led to the discovery of the nucleus. He proposed that:

- The atom has a small, dense, positively charged nucleus at its center.
- Most of the atom is empty space.
- Electrons orbit the nucleus like planets around the sun.





A diagram illustrating Rutherford's gold foil experiment with alpha particles being scattered by the nucleus. Another diagram showing the Rutherford atomic model with a central nucleus and orbiting electrons.

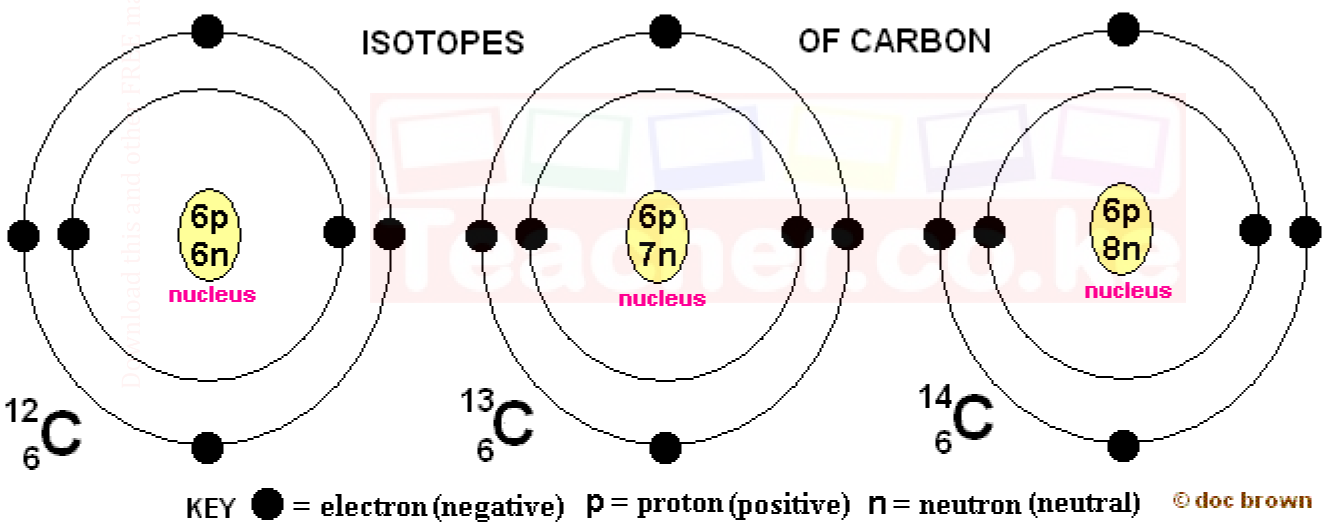
(iii) Atomic Number and Mass Number

- **Atomic Number (Z):** The number of protons in the nucleus of an atom. It uniquely identifies an element. For example, Hydrogen (H) has Z=1, Carbon (C) has Z=6.
- **Mass Number (A):** The total number of protons and neutrons in the nucleus of an atom. Number of neutrons = A - Z. For example, Carbon-12 has A=12 and Z=6, so it has 6 neutrons.

(b) Relative Atomic Mass of Elements

(i) Isotopes

Isotopes are atoms of the same element that have the same number of protons (same atomic number) but different numbers of neutrons (different mass numbers). For example, Carbon-12 (612C), Carbon-13 (613C), and Carbon-14 (614C) are isotopes of carbon.



A diagram showing the nuclei of the three isotopes of carbon, highlighting the different numbers of neutrons.

(ii) Relative Atomic Mass (RAM)

The relative atomic mass of an element is the weighted average of the masses of its naturally occurring isotopes, compared to 1/12th the mass of a carbon-12 atom. It takes into account the abundance of each isotope.

Formula for RAM:

RAM = [(% Abundance of Isotope 1 × Mass of Isotope 1) + (% Abundance of Isotope 2 × Mass of Isotope 2) + ...] / 100

Example:

Chlorine exists as two common isotopes: Chlorine-35 (35Cl) with an abundance of 75.77% and a mass of 35 amu, and Chlorine-37 (37Cl) with an abundance of 24.23% and a mass of 37 amu.

RAM of Chlorine = [(75.77 × 35) + (24.23 × 37)] / 100 = [2651.95 + 896.51] / 100 = 3548.46 / 100 = 35.48 amu

(c) Electron Arrangement using s and p Notation

(i) Energy Levels and Orbitals

Electrons in an atom are arranged in specific energy levels or shells around the nucleus. These energy levels are further divided into sublevels called **orbitals**. Orbitals are regions of space where there is a high probability of finding an electron.

The first energy level (n=1) has one type of orbital called the **s orbital**. The second energy level (n=2) has two types of orbitals: one **s orbital** and three **p orbitals**. The third energy level (n=3) has s, p, and d orbitals, and so on. For the first 20 elements, we mainly focus on s and p orbitals.

- **s orbital:** Spherical in shape. Each s orbital can hold a maximum of **2 electrons**.: A diagram showing the spherical shape of an s orbital.
- **p orbitals:** Dumbbell-shaped. There are three p orbitals in each energy level (except the first one), oriented along the x, y, and z axes (px, py, pz). Each p orbital can hold a maximum of **2 electrons**, so a set of three p orbitals can hold a total of **6 electrons**..

(ii) Order of Filling Electrons (Aufbau Principle)

Electrons fill atomic orbitals in order of increasing energy. For the first 20 elements, the order of filling is approximately:

1s, 2s, 2p, 3s, 3p, 4s

(iii) Writing Electron Arrangement using s and p Notation

The electron arrangement (or electron configuration) shows how electrons are distributed among the orbitals in an atom. The notation indicates the principal energy level (n), the type of orbital (s or p), and the number of electrons in that orbital as a superscript.

Electron Arrangements of the First 20 Elements:

Atomic Number (Z)	Element	Symbol	Electron Arrangement (s and p notation)
1	Hydrogen	H	$1s^1$
2	Helium	He	$1s^2$
3	Lithium	Li	$1s^2\ 2s^1$
4	Beryllium	Be	$1s^2\ 2s^2$
5	Boron	B	$1s^2\ 2s^2\ 2p^1$
6	Carbon	C	$1s^2\ 2s^2\ 2p^2$
7	Nitrogen	N	$1s^2\ 2s^2\ 2p^3$
8	Oxygen	O	$1s^2\ 2s^2\ 2p^4$
9	Fluorine	F	$1s^2\ 2s^2\ 2p^5$
10	Neon	Ne	$1s^2\ 2s^2\ 2p^6$
11	Sodium	Na	$1s^2\ 2s^2\ 2p^6\ 3s^1$
12	Magnesium	Mg	$1s^2\ 2s^2\ 2p^6\ 3s^2$
13	Aluminium	Al	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
14	Silicon	Si	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
15	Phosphorus	P	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
16	Sulfur	S	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
17	Chlorine	Cl	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
18	Argon	Ar	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
19	Potassium	K	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
20	Calcium	Ca	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

(d) Interest in the Study of Structure of the Atom

Understanding the structure of the atom is fundamental to comprehending how elements behave and interact to form molecules and compounds. This knowledge is crucial for advancements in various fields like medicine, materials science, and energy production. By exploring the subatomic world, we gain insights into the very nature of matter.

Activity Suggestion:

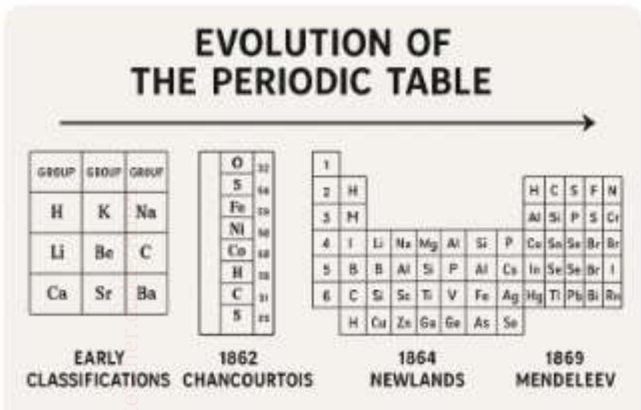
Watch a simulation of the Rutherford Gold Foil experiment and discuss the findings with peers to appreciate the scientific process of understanding atomic structure. You can also engage in simple activities to visualize the filling of electrons into orbitals using diagrams or models.

Sub-Strand 1.3: The Periodic Table

(a) Development of the Periodic Table

The periodic table is a systematic arrangement of chemical elements ordered by their atomic number, electron configuration, and recurring chemical properties.

- **Early Attempts:** Scientists like Johann Wolfgang Döbereiner with his "triads" and John Newlands with his "law of octaves" attempted to classify elements based on their properties and atomic masses.
- **Mendeleev's Periodic Table:** Dmitri Mendeleev is widely credited with developing the first periodic table in 1869. He arranged elements in order of increasing atomic mass and noticed recurring patterns in their properties. He even predicted the existence and properties of undiscovered elements.
- **Modern Periodic Table:** The modern periodic table is arranged by increasing **atomic number** (number of protons), as discovered by Henry Moseley. This arrangement resolved some inconsistencies in Mendeleev's table.



(b) Relating Position in the Periodic Table to Electron Arrangement

The position of an element in the periodic table directly correlates with its electron arrangement:

- **Groups (Vertical Columns):** Elements in the same group have the same number of valence electrons (electrons in the outermost energy level), which leads to similar chemical properties. The group number (for main group elements) often indicates the number of valence electrons.
  - ✓ **Example:** Elements in Group 1 (Alkali Metals) have 1 valence electron (e.g., Sodium: 2, 8, 1 or $1s^2\ 2s^2\ 2p^6\ 3s^1$). Elements in Group 17 (Halogens) have 7 valence electrons (e.g., Chlorine: 2, 8, 7 or $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$).
- **Periods (Horizontal Rows):** Elements in the same period have the same number of occupied electron shells. The period number corresponds to the highest principal energy level occupied by electrons.
  - ✓ **Example:** Elements in Period 2 have electrons in the first and second energy levels (n=1 and n=2). Elements in Period 3 have electrons in the first, second, and third energy levels (n=1, n=2, and n=3).

(c) Ion Formation of Elements using the Periodic Table

Atoms tend to gain, lose, or share electrons to achieve a stable electron configuration, usually resembling that of the nearest noble gas (8 valence electrons, or 2 for Helium). This process leads to the formation of ions:

- **Cations (Positive Ions):** Metals (generally located on the left side and in the middle of the periodic table) tend to lose electrons to achieve a stable configuration. The charge of the cation is equal to the number of electrons lost.
  - ✓ Elements in their elemental form have an oxidation number of 0 .
  - ✓ The sum of oxidation numbers in a neutral compound is zero. For polyatomic ions, the sum equals the charge of the ion.

Deriving Valency and Oxidation Numbers from Electron Arrangement:

By looking at the electron arrangement, you can determine how many electrons an atom needs to gain, lose, or share to achieve stability, thus indicating its valency and likely oxidation states.

- **Example:** Oxygen (O) has an electron arrangement of 2, 6 (or $1s^2\ 2s^2\ 2p^4$). It needs to gain 2 electrons to achieve a stable octet, so its valency is often 2, and its oxidation number in many compounds is -2.
- **Example:** Aluminium (Al) has an electron arrangement of 2, 8, 3 (or $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$). It tends to lose 3 electrons to form Al^{3+}$, so its valency is 3, and its oxidation number is +3.

(f) Deriving Formulae of Compounds

Chemical formulae represent the ratio of atoms of each element in a compound. We can derive them using valencies or oxidation states:

1. Write the symbols of the elements involved.
2. Write their valencies or oxidation numbers.
3. Swap the valencies or oxidation numbers (ignoring the sign for valency).
4. Simplify the ratio if possible.

Examples:

- ✓ **Sodium Chloride:** Sodium (Na) has a valency of 1, Chlorine (Cl) has a valency of 1. Swapping gives Na_1Cl_1, so the formula is **NaCl**.
- ✓ **Magnesium Oxide:** Magnesium (Mg) has a valency of 2, Oxygen (O) has a valency of 2. Swapping gives Mg_2O_2, simplifying the ratio gives **MgO**.
- ✓ **Aluminium Oxide:** Aluminium (Al) has a valency of 3, Oxygen (O) has a valency of 2. Swapping gives Al_2O_3, so the formula is **Al_2O_3**.

(g) Writing Balanced Chemical Equations for Simple Chemical Reactions

A balanced chemical equation represents a chemical reaction with the same number of atoms of each element on both the reactant and product sides.

1. Write the unbalanced equation with the correct chemical formulae of reactants and products.
2. Count the number of atoms of each element on both sides.
3. Adjust the coefficients (numbers in front of the formulae) to balance the number of atoms of each element. Start with elements that appear in only one reactant and one product.

Example: Reaction between Magnesium and Oxygen to form Magnesium Oxide.

Unbalanced equation: $Mg(s) + O_2(g) \rightarrow MgO(s)$

- ✓ Mg: 1 on left, 1 on right (Balanced)
- ✓ O: 2 on left, 1 on right (Unbalanced)

Balance Oxygen: $Mg(s) + O_2(g) \rightarrow 2MgO(s)$

- ✓ Mg: 1 on left, 2 on right (Unbalanced)
- ✓ O: 2 on left, 2 on right (Balanced)

Balance Magnesium: $\mathbf{2}Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Balanced equation: $\mathbf{2Mg(s) + O_2(g) \rightarrow 2MgO(s)}$

Suggestion: Examples of simple chemical reactions with corresponding balanced chemical equations, possibly with molecular models illustrating the rearrangement of atoms.

(h) Appreciation of Electron Arrangement in the Development of the Periodic Table

The periodic table is a testament to the fundamental role of electron arrangement in determining the chemical properties of elements. The recurring patterns observed in the periodic table directly reflect the periodic filling of electron shells and subshells. This understanding allows us to predict the behavior of elements and design new materials and technologies. The arrangement of elements based on their electron configurations provides a logical and powerful framework for studying chemistry.

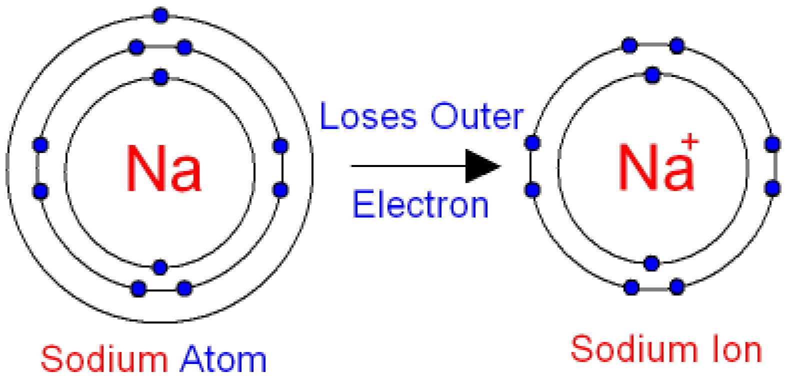
Sub-Strand 1.4: Chemical Bonding (24 Lessons)

(a) Illustrating Bond Types in Elements, Molecules, and Compounds

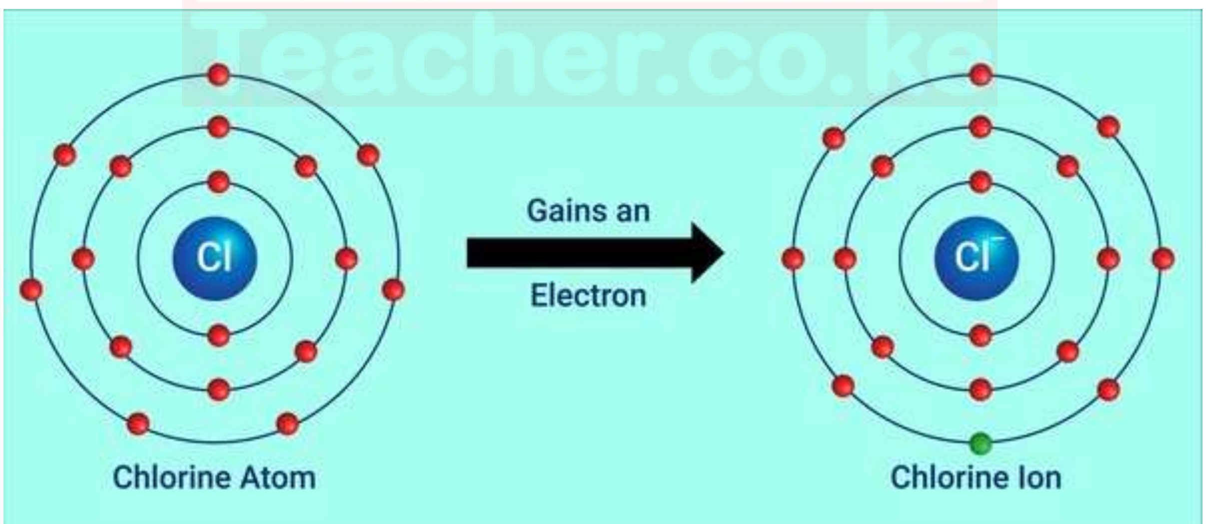
Atoms combine to form molecules and compounds through chemical bonds, striving to achieve a stable electron configuration (usually like that of a noble gas). The valence electrons (outermost electrons) play a crucial role in this process.

Types of Chemical Bonds:

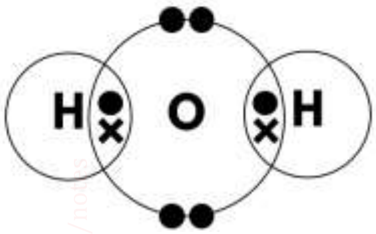
- **Ionic Bond:** Formed by the complete transfer of one or more electrons from one atom (usually a metal) to another (usually a nonmetal), resulting in the formation of oppositely charged ions (cations and anions) that are attracted to each other by electrostatic forces.
  - ✓ **Example:** Sodium Chloride (NaCl). Sodium (Na) loses one electron to become a sodium ion,



Chlorine (Cl) gains one electron to become a Chlorine ion.



- **Covalent Bond:** Formed by the sharing of one or more pairs of electrons between two nonmetal atoms. This sharing allows both atoms to achieve a stable electron configuration.
  - ✓ **Example:** Water (H_2O). Oxygen shares electrons with two hydrogen atoms, forming covalent bonds.



- **Dative Covalent Bond (Coordinate Bond):** A type of covalent bond where one of the bonded atoms provides both electrons for the shared pair.
    - ✓ **Example:** Ammonium ion (NH_4+).
- **Water (H_2O):** Covalent bonding within the molecule and hydrogen bonding between molecules result in a relatively higher boiling point for a simple molecular compound.
- **Diamond (Carbon):** Giant atomic structure with strong covalent bonds in a tetrahedral network results in extreme hardness and a very high melting point; it does not conduct electricity.
- **Graphite (Carbon):** Giant atomic structure with carbon atoms arranged in layers of hexagons with strong covalent bonds within the layers but weak Van der Waals forces between layers. This allows layers to slide, making graphite soft and a good lubricant; it also conducts electricity due to delocalized electrons within the layers. **Silicon Dioxide (SiO_2 - Quartz):** Giant atomic structure with silicon and oxygen atoms covalently bonded in a network, resulting in hardness and a high melting point. **Aluminium (Al):** Metallic bonding with delocalized electrons accounts for its good electrical and thermal conductivity, malleability, and ductility

(c) Relating Bond Types and Resultant Structures to Uses

The properties resulting from different bond types and structures determine the uses of substances:

- **Ionic Compounds (e.g., NaCl - Table Salt):** Used as seasoning and food preservation due to its crystalline structure and solubility.
- **Simple Covalent Molecules (e.g., Water - H_2O):** Essential for life as a solvent and reactant due to its polarity and ability to form hydrogen bonds.
- **Giant Atomic/Covalent Structures:**
    - **Diamond:** Used in cutting tools and abrasives due to its extreme hardness. **Graphite:** Used as a lubricant and in pencil leads due to its layered structure allowing layers to slide. It is also used as an electrode due to its conductivity. **Silicon Dioxide (Quartz):** Used in making glass (windows, containers) due to its transparency and hardness. Also used in electronics.

- **Metallic Compounds (Metals):**
  - **Aluminium:** Used in aircraft construction, cans, and electrical wires due to its low density, strength, and conductivity. **Copper:** Used extensively in electrical wiring and plumbing due to its high conductivity and malleability. **Iron:** Used in construction and making tools due to its strength and availability (often in alloys like steel).

(d) Appreciation of Uses Based on Bond Types and Structures

Understanding chemical bonding allows us to explain why different substances have different properties and why they are suitable for specific applications. From the hardness of diamond enabling cutting to the conductivity of copper facilitating electricity flow, the microscopic arrangement of atoms and the forces holding them together dictate the macroscopic world around us. This knowledge is crucial for innovation and the development of new materials with tailored properties for various needs in our daily lives.

Project Idea:

Model bonding in selected molecules or compounds like NaCl, SiO_2, graphite, and diamond using locally available materials such as balls, sticks, or clay to visualize the different structures and bonding arrangements.

Sub-Strand 1.5: Periodicity

(a) Trends in Physical Properties of Elements in Groups

Physical properties of elements show trends as you move down a group in the periodic table. We will focus on Groups I, II, VII, and VIII.

(i) Group I: Alkali Metals (Lithium, Sodium, Potassium, etc.)

- ✓ **Appearance:** Silvery-white, soft, shiny metals. React readily with air and tarnish quickly.
- ✓ **Atomic and Ionic Size:** Increases down the group due to the addition of electron shells.
- ✓ **Ionization Energy:** Decreases down the group because the outermost electron is further from the nucleus and easier to remove.
- ✓ **Electron Affinity:** Generally low and becomes less negative down the group (tendency to gain electrons decreases).
- ✓ **Electronegativity:** Decreases down the group as the attraction for electrons in a bond weakens.

- ✓ **Melting and Boiling Points:** Decrease down the group due to weaker metallic bonding as atomic size increases.
- ✓ **Ductility and Malleability:** Soft metals that can be easily cut, shaped, and drawn into wires.
- ✓ **Electrical Conductivity:** Good conductors of electricity due to the presence of one delocalized electron per atom.

(ii) Group II: Alkaline Earth Metals (Beryllium, Magnesium, Calcium, etc.)

- ✓ **Appearance:** Silvery-white metals, harder and denser than alkali metals. React with air but less vigorously.
- ✓ **Atomic and Ionic Size:** Increases down the group.
- ✓ **Ionization Energy:** Decreases down the group (though higher than Group I elements due to two valence electrons and smaller atomic size).
- ✓ **Electron Affinity:** Generally low.
- ✓ **Electronegativity:** Decreases down the group.
- ✓ **Melting and Boiling Points:** Decrease down the group (generally higher than Group I).
- ✓ **Ductility and Malleability:** Malleable and ductile, but less so than alkali metals.
- ✓ **Electrical Conductivity:** Good conductors of electricity (with two delocalized electrons per atom).

(iii) Group VII: Halogens (Fluorine, Chlorine, Bromine, Iodine)

- **Appearance:** Nonmetals with varying physical states and colors: Fluorine (pale yellow gas), Chlorine (greenish-yellow gas), Bromine (red-brown liquid), Iodine (dark violet solid). Smell is pungent and often irritating.
- **Atomic and Ionic Size:** Increases down the group.
- **Ionization Energy:** Decreases down the group as the outermost electron is further from the nucleus.
- **Electron Affinity:** High and negative (readily accept one electron to form -1 ion), generally decreases down the group.
- **Electronegativity:** Highest among the elements in their respective periods, decreases down the group (Fluorine is the most electronegative element).
- **Melting and Boiling Points:** Increase down the group due to increasing strength of Van der Waals forces as molecular size increases.
- **Ductility and Malleability:** Brittle nonmetals, do not exhibit ductility or malleability.
- **Electrical Conductivity:** Poor conductors of electricity as they lack free mobile electrons.

(iv) Group VIII: Noble Gases (Helium, Neon, Argon, Krypton, Xenon)

- **Appearance:** Colorless, odorless, and tasteless gases under normal conditions.
- **Atomic Size:** Increases down the group.
- **Ionization Energy:** Very high because they have a stable, full outer electron shell, making it very difficult to remove an electron. Decreases slightly down the group due to increasing atomic size.
- **Electron Affinity:** Very low (close to zero) as they have no tendency to gain electrons.
- **Electronegativity:** Considered to be zero or very low as they generally do not form bonds.
- **Melting and Boiling Points:** Very low, increasing down the group due to increasing strength of London dispersion forces with increasing atomic size.
- **Ductility and Malleability:** Do not exhibit these properties as they are gases.
- **Electrical Conductivity:** Poor conductors of electricity under normal conditions.

(b) Chemical Properties of Elements in Groups

Chemical properties are related to how elements react.

(i) Group I: Alkali Metals

- **Reaction with Oxygen:** Readily react with oxygen in the air to form metal oxides. The type of oxide formed can vary (e.g., Na can form Na_2O, Na_2O_2, NaO_2). **Experiment Suggestion:** React a small piece of sodium metal with air on a watch glass and observe the changes.
- **Reaction with Chlorine:** React vigorously with chlorine gas to form white crystalline ionic chlorides (metal halides). **Experiment Suggestion:** Gently heat sodium metal in a deflagrating spoon and introduce it into a gas jar filled with chlorine gas (prepared carefully in a fume hood). Observe the reaction.
- **Reaction with Cold Water:** React vigorously with cold water to produce hydrogen gas and a metal hydroxide solution (alkaline). The reactivity increases down the group.

  Experiment Suggestion:

  Add small pieces of lithium, sodium, and potassium separately to troughs of water containing phenolphthalein indicator.

  Observe the rate of reaction and any color changes.

  Collect the gas produced by sodium and test it with a burning splint (should produce a "pop" sound confirming hydrogen).

- **Reaction with Steam:** React with steam to produce hydrogen gas and a metal oxide.
- **Reaction with Dilute Acids:** React readily with dilute acids to produce hydrogen gas and a metal salt.

(ii) Group II: Alkaline Earth Metals

- **Reaction with Oxygen:** Burn in oxygen to form metal oxides (e.g., MgO, CaO). Reactivity increases down the group.
- **Reaction with Chlorine:** React with chlorine gas upon heating to form metal chlorides (e.g., $MgCl_2$, $CaCl_2$).
- **Reaction with Cold Water:** React with cold water to form hydrogen gas and a metal hydroxide. Reactivity increases down the group (Beryllium does not react, Magnesium reacts very slowly, Calcium reacts more readily).
- **Reaction with Steam:** React with steam to form hydrogen gas and a metal oxide (Magnesium reacts with steam).
- **Reaction with Dilute Acids:** React with dilute acids to produce hydrogen gas and a metal salt.

(iii) Group VII: Halogens

- **Reaction with Water:** Chlorine reacts with water to form hydrochloric acid (HCl) and hypochlorous acid ($HOCl$). Bromine and iodine are less soluble in water and react to a lesser extent.

  Experiment Suggestion:

  Bubble chlorine gas (prepared carefully in a fume hood) into water and test the acidity of the solution using litmus paper.

- **Reaction with Metals:** React with many metals to form metal halides (e.g., reaction of chlorine with iron wool to form iron(III) chloride).

  Experiment Suggestion:

  Heat iron wool in a combustion tube and pass dry chlorine gas over it.

  Observe the formation of brown solid iron(III) chloride.

- **Displacement Reactions:** A more reactive halogen can displace a less reactive halogen from its salt solution (Reactivity order: Fluorine > Chlorine > Bromine > Iodine).

Experiment Suggestion:

Add chlorine water to potassium bromide solution and observe the formation of orange-brown bromine.

Then add chlorine water to potassium iodide solution and observe the formation of dark solid iodine.

- **Bleaching Action:** Chlorine (and to some extent other halogens) can act as a bleaching agent due to the formation of hypochlorous acid in water, which decomposes to release oxygen atoms that cause bleaching.

(iv) Group VIII: Noble Gases

- Generally unreactive due to their stable electron configurations. However, heavier noble gases like Xenon can form some compounds with highly electronegative elements like Fluorine under specific conditions.

(c) Trends in Properties Across Period 3 (Sodium to Argon)

Trends across a period generally involve an increase in the effective nuclear charge experienced by the valence electrons as the number of protons increases while electrons are added to the same energy level.

- **Atomic Size:** Decreases across the period. The increasing nuclear charge pulls the electrons closer to the nucleus..
- **Ionization Energy:** Generally increases across the period. It becomes harder to remove an electron due to the increasing nuclear attraction. There are slight exceptions at Magnesium and Sulfur due to stable electron configurations.
- **Electron Affinity:** Generally increases across the period (becomes more negative), except for Group 2 and Group 18 elements which have stable configurations.
- **Electronegativity:** Increases across the period as the tendency to attract electrons in a bond increases due to the increasing nuclear charge.
- **Melting and Boiling Points:** Generally increase from Sodium to Silicon (due to increasing strength of metallic/covalent bonds) and then decrease from Phosphorus to Argon (due to weaker intermolecular forces in simple molecular and atomic structures).
- **Electrical Conductivity:** Decreases across the period from metals (Na, Mg, Al) to metalloid (Si) to nonmetals (P, S, Cl, Ar). Silicon is a semiconductor.

(d) Applications of Elements of the Periodic Table (Groups I, II, VII, VIII)

- **Group I (Alkali Metals):**

- ✓ **Lithium:** Used in batteries for electronics and electric vehicles, in some medications for mental health.
- ✓ **Sodium:** Used in streetlights (sodium vapor lamps), as a heat transfer medium in some nuclear reactors, and in the production of many chemicals.
- ✓ **Potassium:** Essential nutrient for plants (used in fertilizers), and plays a vital role in nerve function in animals. Potassium hydroxide is used in making soaps.

- **Group II (Alkaline Earth Metals):**
  - ✓ **Magnesium:** Used in lightweight alloys for aircraft and cars, in flares and fireworks, and as a dietary supplement.
  - ✓ **Calcium:** Essential for bones and teeth, used in cement and plaster of Paris, and in agriculture to neutralize acidic soils.
- **Group VII (Halogens):**
  - ✓ **Chlorine:** Used to disinfect water supplies and swimming pools, as a bleaching agent in the paper and textile industries, and in the production of plastics like PVC.
  - ✓ **Iodine:** Used as an antiseptic in wound cleaning solutions, and as an essential nutrient for the thyroid gland (added to table salt).
- **Group VIII (Noble Gases):**
  - ✓ **Helium:** Used in balloons and airships due to its low density, as a coolant in superconducting magnets (like those in MRI machines), and in deep-sea diving mixtures.
  - ✓ **Neon:** Used in advertising signs because it glows red when electricity is passed through it.
  - ✓ **Argon:** Used as an inert atmosphere for welding and in incandescent light bulbs to prevent the filament from burning out.

(e) Appreciation of Applications of Various Elements

The diverse properties of elements, arising from their electronic structures and their position in the periodic table, make them indispensable for a wide range of applications that improve our quality of life, from providing light and energy to ensuring our health and safety. Understanding these periodic trends allows scientists to predict the behavior of known and undiscovered elements and to develop new technologies and materials.

STRAND 2.0: PHYSICAL CHEMISTRY

Sub-Strand 2.1: Acids and Bases

(a) Characteristics of Acids and Bases in Aqueous Solutions

Acids:

- Have a **sour taste** (though you should *never* taste chemicals in the lab!).
- Are typically **corrosive**.
- Turn **blue litmus paper red**.
- Produce hydrogen ions

Bases:

- Have a **bitter taste** (again, *never* taste lab chemicals!).
- Feel **slippery to the touch**.
- Turn **red litmus paper blue**.
- Produce hydroxide ions

(b) Chemical Properties of Acids and Bases

Chemical Properties of Acids:

- **Reaction with Metals:** Acids react with most reactive metals to produce a **salt** and **hydrogen gas**.
  - ✓ **General Equation:** Acid + Metal → Salt + Hydrogen gas
  - ✓ **Example:**

Hydrochloric acid + Zinc → Zinc chloride + Hydrogen gas

$2HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)\uparrow$

Experiment Suggestion:

- React a piece of magnesium ribbon with dilute hydrochloric acid in a test tube and collect the gas produced.
- Test the gas with a burning splint (it should make a "pop" sound, confirming hydrogen).

- **Reaction with Metal Carbonates and Hydrogen Carbonates:** Acids react with metal carbonates and hydrogen carbonates to produce a **salt**, **water**, and **carbon dioxide gas**.
  - ✓ **General Equation:**

Acid + Metal Carbonate → Salt + Water + Carbon Dioxide gas

  - ✓ **Example:**

Sulfuric(VI) acid + Sodium carbonate → Sodium sulfate + Water + Carbon dioxide gas

$H_2SO_4(aq) + Na_2CO_3(s) \rightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)\uparrow$

  - ✓ **General Equation:** Acid + Metal Hydrogen Carbonate → Salt + Water + Carbon Dioxide gas
  - ✓ **Example:**

Hydrochloric acid + Sodium hydrogen carbonate → Sodium chloride + Water + Carbon dioxide gas

$HCl(aq) + NaHCO_3(s) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)\uparrow$

Experiment Suggestion:

Add dilute hydrochloric acid to a sample of sodium carbonate in a test tube.

Bubble the gas produced through limewater (calcium hydroxide solution).

The limewater should turn milky, confirming the presence of carbon dioxide.

- **Reaction with Metal Oxides:** Acids react with basic metal oxides to form a **salt** and **water** (neutralization reaction).
  - ✓ **General Equation:** Acid + Metal Oxide → Salt + Water
  - ✓ **Example:** Hydrochloric acid + Copper(II) oxide → Copper(II) chloride + Water
    - $2HCl(aq) + CuO(s) \rightarrow CuCl_2(aq) + H_2O(l)$
- **Reaction with Metal Hydroxides (Alkalis):** Acids react with metal hydroxides (bases) to form a **salt** and **water** (neutralization reaction).
  - ✓ **General Equation:** Acid + Metal Hydroxide → Salt + Water
  - ✓ **Example:** Sulfuric(VI) acid + Sodium hydroxide → Sodium sulfate + Water
    - $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

Chemical Properties of Bases:

- **Reaction with Acids:** Bases react with acids to form a **salt** and **water** (neutralization reaction). (See examples above).
- **Reaction with Ammonium Salts:** Some bases, especially alkalis, react with ammonium salts upon heating to produce a **salt**, **water**, and **ammonia gas**.
  - ✓ **General Equation:** Base + Ammonium Salt → Salt + Water + Ammonia gas
  - ✓ **Example:** Sodium hydroxide + Ammonium chloride → Sodium chloride + Water + Ammonia gas
    - $NaOH(aq) + NH_4Cl(s) \rightarrow NaCl(aq) + H_2O(l) + NH_3(g)\uparrow$

Experiment Suggestion:

Heat a mixture of ammonium chloride and sodium hydroxide in a test tube.

Test the gas produced with a piece of moist red litmus paper.

It should turn blue, confirming the presence of ammonia.

(c) Classification of Acids and Bases into Strong and Weak using Universal Indicator

The strength of an acid or base depends on the extent to which it dissociates or ionizes in water to produce H^+ or OH^- ions, respectively.

- **Strong Acids:** Dissociate almost completely in water, producing a high concentration of H^+ ions. Examples include hydrochloric acid (HCl), sulfuric(VI) acid (H_2SO_4), and nitric acid (HNO_3). .
- **Weak Acids:** Dissociate only partially in water, producing a low concentration of H^+ ions. Examples include ethanoic acid (CH_3COOH) and carbonic acid (H_2CO_3). .
- **Strong Bases:** Dissociate or ionize completely in water, producing a high concentration of OH^- ions. Examples include sodium hydroxide (NaOH), potassium hydroxide (KOH).
- **Weak Bases:** Dissociate or ionize only partially in water, producing a low concentration of OH$^-$ ions. Example includes ammonia (NH_3).

Universal Indicator and pH Scale:

A **universal indicator** is a mixture of several indicators that changes color gradually over a wide range of pH values. The **pH scale** is a numerical scale from 0 to 14 that indicates the acidity or alkalinity of a solution:

- **pH < 7:** Acidic solution (lower pH means stronger acid)
- **pH = 7:** Neutral solution (e.g., pure water)
- **pH > 7:** Basic or alkaline solution (higher pH means stronger base)

Electrical Conductivity:

Strong acids and bases are **good conductors of electricity** in aqueous solutions because they produce a high concentration of mobile ions. Weak acids and bases are **poor conductors of electricity** because they produce a low concentration of mobile ions.

Experiment Suggestion:

a) Prepare aqueous solutions of strong acids (e.g., HCl, H_2SO_4), weak acids (e.g., CH_3COOH), strong bases (e.g., NaOH), and weak bases (e.g., NH_3) of the same concentration.
b) Dip the electrodes of a conductivity meter into each solution and observe the brightness of the bulb (or the reading on the meter) to compare their electrical conductivity.
c) Test the pH of each solution using a universal indicator paper or solution and correlate it with the conductivity.

(d) Outline the Uses of Acids and Bases in Day-to-Day Life

Uses of Acids:

- **Hydrochloric acid (HCl):** Used in the stomach for digestion, in industries for cleaning metals (pickling), and in the production of dyes and plastics
- **Sulfuric(VI) acid (H_2SO_4):** Used in the production of fertilizers, detergents, dyes, and in car batteries. It is also a strong dehydrating agent.
- **Ethanoic acid (CH_3COOH) (Vinegar):** Used in food preservation and as a flavoring agent in cooking.
- **Citric acid:** Found in citrus fruits (lemons, oranges) and used as a flavoring and preservative in food and drinks.

Uses of Bases:

- **Sodium hydroxide (NaOH) (Caustic Soda):** Used in the manufacture of soap and detergents, in the paper industry, and as a drain cleaner.
- **Calcium hydroxide ($Ca(OH)_2$) (Slaked Lime):** Used in agriculture to neutralize acidic soils, in the production of cement and mortar, and in water treatment.
- **Ammonia (NH_3):** Used in the production of fertilizers, as a cleaning agent, and in the manufacture of nitric acid.
- **Magnesium hydroxide ($Mg(OH)_2$):** Used in some antacids to neutralize excess stomach acid.
- **Sodium carbonate (Na_2CO_3) (Washing Soda):** Used in laundry detergents and in the manufacture of glass.

(e) Appreciation of the Uses of Acids and Bases

Acids and bases are essential chemicals with a wide range of applications that impact our daily lives significantly. From industrial processes and agriculture to household cleaning and even within our own bodies, their unique properties make them invaluable in various fields. Understanding their characteristics and safe handling is crucial for their effective and responsible use.

Sub-Strand 2.2: Introduction to Salts

(a) Classification of Salts Based on Their Properties

A salt is an ionic compound formed when the hydrogen ions (H^+) of an acid are fully or partially replaced by a metal or ammonium ion.

Classification:

Normal (Neutral) Salts

Formed by the complete replacement of all replaceable hydrogen ions of an acid by a metal or ammonium ion.

Result in a neutral solution when dissolved in water (pH ≈ 7).

Example: Sodium chloride (NaCl), potassium sulfate (K_2SO_4).

Acidic Salts (Hydrogen Salts):

- Formed by the partial replacement of the replaceable hydrogen ions of a polyprotic acid (an acid with more than one replaceable H^+) by a metal or ammonium ion.
- Contain remaining replaceable hydrogen ions, making them capable of reacting with bases.
- Result in an acidic solution when dissolved in water (pH < 7).
- Example: Sodium hydrogen sulfate ($NaHSO_4$), sodium hydrogen carbonate ($NaHCO_3$).

Basic Salts (Hydroxide Salts):

- Contain hydroxide ions (OH^-) along with the metal cation and the anion of the acid.
- Formed by the incomplete neutralization of a polyacidic base (a base with more than one OH^-) by an acid.
- Examples of polyacidic bases are aluminum hydroxide ($Al(OH)_3$) and iron(III) hydroxide ($Fe(OH)_3$).

- Result in a basic (alkaline) solution when dissolved in water (pH > 7).
- Example: Magnesium hydroxide chloride ($Mg(OH)Cl$), lead(II) hydroxide nitrate ($Pb(OH)NO_3$).

***Double Salts:**

- Formed by the combination of two simple salts with different cations but the same anion, or vice versa, which crystallize together in a fixed ratio.
- Dissociate into their constituent ions when dissolved in water.
- **Example:** Potassium aluminum sulfate ($KAl(SO_4)_2 \cdot 12H_2O$) - Alum, Mohr's salt ($(NH_4)_2Fe(SO_4)_2 \cdot 6H_2O$).
- Preparation of Salts Using Appropriate Methods in the Laboratory**

General Principle:

The method of preparation depends on the solubility of the desired salt.

Preparation of Soluble Salts:

Reaction between an Acid and a Reactive Metal:

- ✓ Applicable for metals above hydrogen in the reactivity series.
  General equation:
  Acid + Metal → Salt + Hydrogen gas
  Example: $2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)\uparrow$

  Experiment Suggestion:

React magnesium ribbon with dilute hydrochloric acid in a beaker.

Evaporate the solution to obtain crystals of magnesium chloride.

Safety Note:
Hydrogen gas is flammable, so keep flames away.

- ✓ Reaction between an Acid and a Base (Neutralization):

General equation:

Acid + Base → Salt + Water

- ✓ Example: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

Experiment Suggestion:

Titrate a known volume of sodium hydroxide solution with hydrochloric acid using phenolphthalein as an indicator. Evaporate the resulting solution to obtain crystals of sodium chloride.

Filtrate: The liquid which has been filtered.

Reaction between an Acid and a Metal Carbonate or Hydrogen Carbonate:

General equation:

Acid + Metal Carbonate → Salt + Water + Carbon Dioxide gas

Example: $H_2SO_4(aq) + CuCO_3(s) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)\uparrow$

Experiment Suggestion:

- ✓ Add sulfuric(VI) acid to copper(II) carbonate in a beaker until no more reaction occurs.
- ✓ Filter the mixture to remove excess copper(II) carbonate, then evaporate the filtrate to obtain crystals of copper(II) sulfate.

Direct Synthesis (Combination)

Applicable for some salts, especially halides and sulfides of less reactive metals.

General equation: Metal + Nonmetal → Salt

Example: $Fe(s) + S(s) \rightarrow FeS(s)$ (upon heating)

Experiment Suggestion:

Heat iron filings and sulfur powder in a crucible until they react to form iron(II) sulfide.

Safety Note: Perform this in a fume hood, as sulfur fumes can be harmful.

Preparation of Insoluble Salts:

Precipitation Reaction:

By mixing aqueous solutions of two soluble salts containing the desired cation and anion.

General equation:

Soluble Salt 1 + Soluble Salt 2 → Insoluble Salt (Precipitate) + Soluble Salt 3

Example:

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s)\downarrow + NaNO_3(aq)$

Ionic equation:

$Ag^+ (aq) + Cl^- (aq) \rightarrow AgCl(s)$

Experiment Suggestion:

a) Mix aqueous solutions of silver nitrate and sodium chloride.
b) Observe the formation of a white precipitate of silver chloride.
c) Filter the precipitate, wash it with distilled water, and dry it to obtain pure silver chloride.

(c) Behaviour of Salts When Exposed to Air

Some salts interact with atmospheric moisture in different ways:

- **Hygroscopic Salts:** These salts absorb moisture from the air but do not dissolve in it to form a solution. They become damp or sticky.
  - **Example:** Anhydrous copper(II) sulfate ($CuSO_4$), Calcium chloride ($CaCl_2$), Sodium hydroxide (NaOH).
- **Deliquescent Salts:** These salts absorb a large amount of moisture from the air and dissolve in it to form a saturated solution. They eventually turn into a liquid.

Example: Sodium hydroxide (NaOH), Potassium hydroxide (KOH), Magnesium chloride ($MgCl_2$).

- **Efflorescent Salts:** These are hydrated salts that lose their water of crystallization when exposed to dry air, forming a powdery layer on their surface.

Example: Washing soda ($Na_2CO_3 \cdot 10H_2O$), Hydrated copper(II) sulfate ($CuSO_4 \cdot 5H_2O$).

Experiment Suggestion:

Place samples of different salts (e.g., anhydrous copper(II) sulfate, sodium hydroxide pellets, hydrated sodium carbonate crystals) on separate watch glasses and leave them exposed to the air for a day or two. Observe and record any changes in their appearance.

(d) Applications of Salts in Day-to-Day Life

Salts have numerous applications in various fields:

- ✓ **Agriculture:** Many salts are used as **inorganic fertilizers** to provide essential nutrients for plant growth. Examples include ammonium sulfate ($(NH_4)_2SO_4$), potassium nitrate (KNO_3), and sodium nitrate ($NaNO_3$).
- ✓ **Food Industry: Sodium chloride (NaCl)** is common **table salt** used for seasoning and as a food preservative. Other salts are used as food additives and preservatives.
- ✓ **Medicine:** Various salts are used in medicines. For example, **magnesium sulfate ($MgSO_4$)** (Epsom salt) is used for muscle relaxation, and **sodium chloride (NaCl)** is used in saline solutions.
- ✓ **Paper Industry:** Salts like **sodium sulfate (Na_2SO_4)** are used in the pulping process.
- ✓ **Paints Industry:** Some salts are used as pigments in paints.
- ✓ **Glass Industry:** Salts like **sodium carbonate (Na_2CO_3)** are used in the manufacture of glass.
- ✓ **Laundry: Sodium carbonate (Na_2CO_3)** (washing soda) is used in laundry detergents to soften water and help remove stains.

(e) Environmental Effects of Inorganic Fertilizers and Mitigation Measures

While inorganic fertilizers increase crop yields, their overuse can have negative environmental impacts:

- ✓ **Water Pollution (Eutrophication):** Excess nitrates and phosphates from fertilizers can leach into rivers and lakes, causing excessive growth of algae (algal blooms). When these algae die and decompose, they deplete the oxygen in the water, harming or killing aquatic organisms.
- ✓ **Soil Pollution:** Overuse of some fertilizers can lead to the build-up of harmful substances in the soil and can affect soil pH and microbial activity.
- ✓ **Air Pollution:** The production and use of some nitrogen-based fertilizers can release greenhouse gases like nitrous oxide (N_2O), contributing to air pollution and climate change.

Mitigation Measures for Sustainable Use:

- ✓ **Use fertilizers judiciously:** Apply fertilizers in the right amounts and at the right time based on soil testing and crop requirements.
- ✓ **Adopt sustainable farming practices:** Crop rotation, cover cropping, and no-till farming can improve soil health and reduce the need for excessive fertilizers.

- ✓ **Use organic fertilizers:** Compost, manure, and other organic materials can provide nutrients to the soil while improving its structure and water retention capacity.
- ✓ **Improve drainage:** Proper drainage can reduce the leaching of fertilizers into water bodies.
- ✓ **Develop slow-release fertilizers:** These fertilizers release nutrients gradually, reducing the risk of nutrient runoff.
- ✓ **Precision agriculture:** Use technology like GPS and sensors to apply fertilizers only where and when needed.