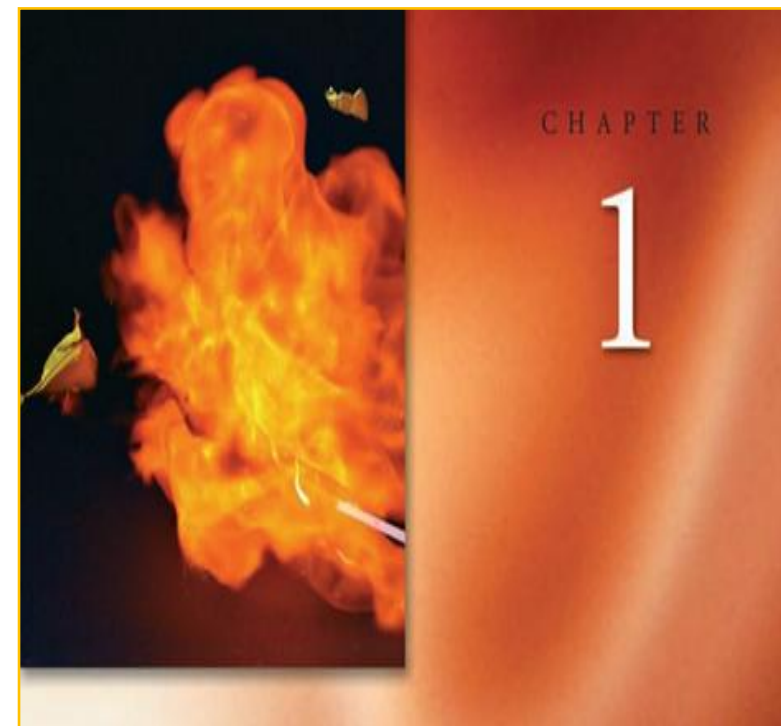


Chemistry and Matter

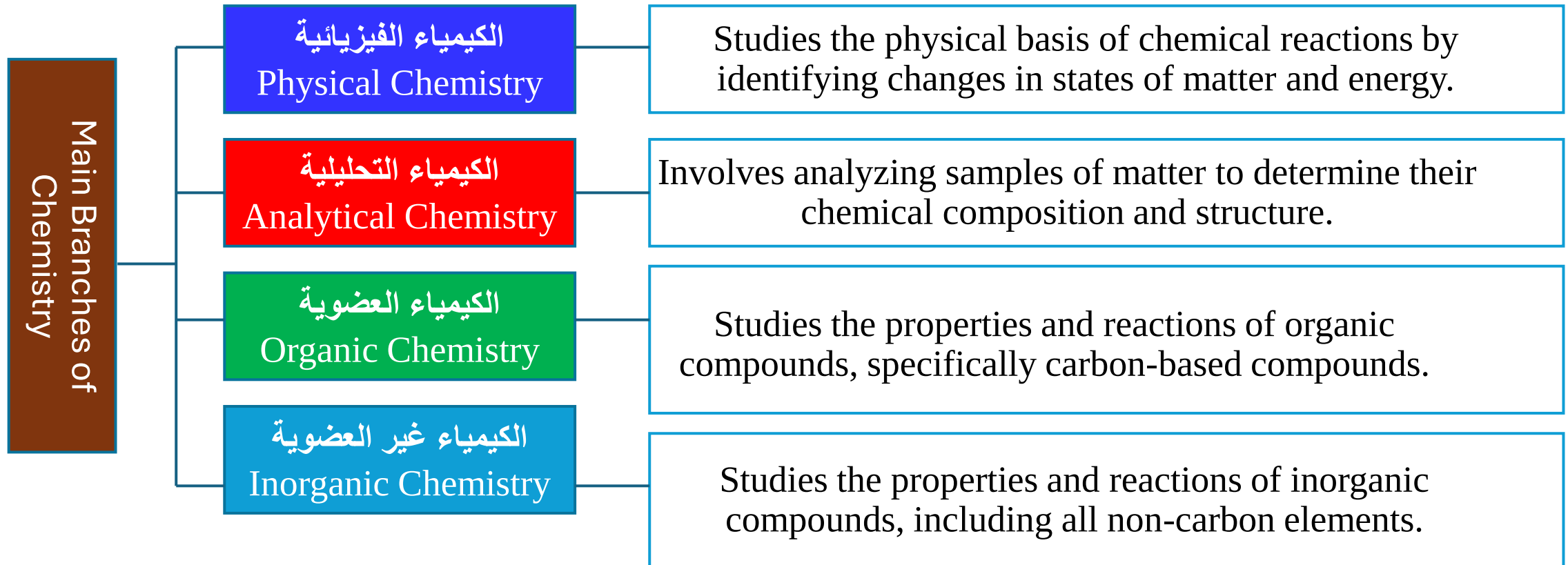
Contents

- **Introduction to the Study of Chemistry**
 - Chemistry and its branches
 - The scientific method
 - The difference between theory and scientific law
- **Matter and Its Properties**
 - Definition and classification of matter
 - States of matter
 - Physical and chemical properties of matter
 - Physical and chemical changes of matter
 - Forms of matter
 - Elements, compounds, and mixtures
 - Physical and chemical methods for separating matter



Introduction to the Study of Chemistry

Chemistry: An experimental science concerned with studying matter in terms of its composition, properties, reactions, and changes occurring within it.

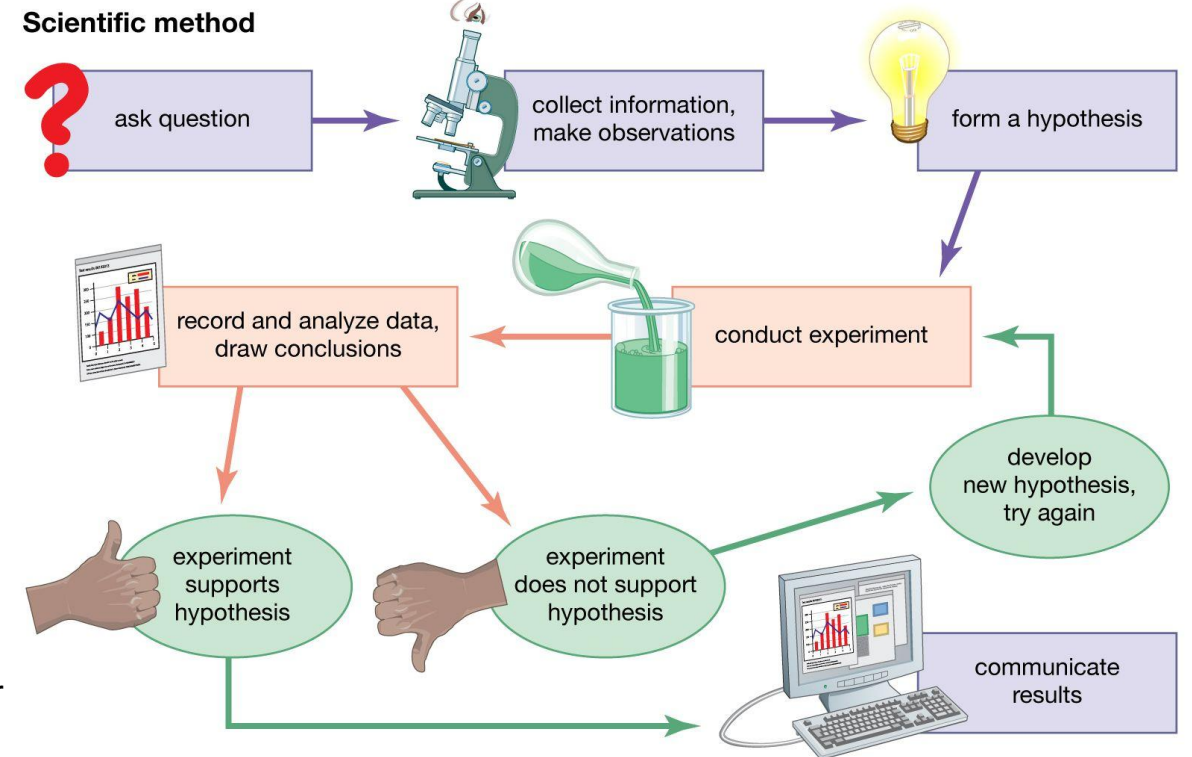


➤ The scientific method is used to study natural sciences such as chemistry, physics, and biology.

Introduction to the Study of Chemistry

The Scientific Method: A scientific approach used to answer questions aimed at explaining natural phenomena.

- Identify a problem or question to be answered through scientific investigation.
- Gather data and observe the problem to gain a better understanding of the issue. This often involves background research.
- Propose a testable explanation or prediction that addresses the question.
- Design and carry out experiments to test the hypothesis. This involves controlling variables and collecting data.
- Evaluate the experimental results to determine if they support or contradict the hypothesis.
- If the experiment **supports the hypothesis**, the findings can be validated and communicated.
- If the experiment **does not support the hypothesis**, modify or develop a new hypothesis and repeat the process.
- Share the findings with others through presentations, publications, or reports.



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Difference Between a Theory and a Scientific Law

- **Theory** (النظرية): A scientific assumption based on logical predictions supported by experiments. It is not an absolute fact.
- **Scientific Law** (القانون العلمي): A theory that has been repeatedly validated and universally accepted as true over a long time.

Matter and Its Properties

Matter is anything that occupies space and has mass. It exists in various forms and can be classified based on its source and origin.

Classification of Matter

- **Natural Matter:**

- Found in the natural environment without human intervention.
- Examples: Apple, Sand, Cotton.

- **Synthetic Matter:**

- Created or modified by humans in laboratories or factories.
- Examples: Plastic, Ibuprofen, Biodegradable Plastics.



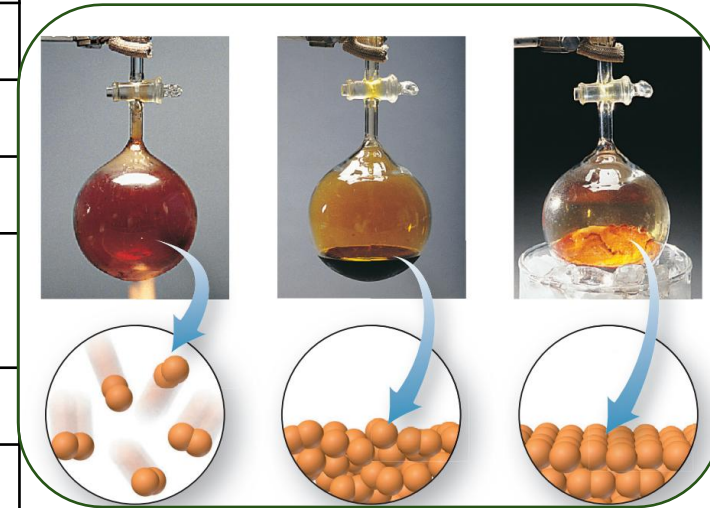
Matter and Its Properties

❑ The States of Matter

- Matter exists in three physical states: solid, liquid, and gas.

Comparison Between States of Matter

Property	Solid State	Liquid State	Gaseous State
Volume	Fixed	Fixed	Not fixed
Shape	Fixed	Variable	Variable
Strength of Bonds	Very strong	Medium	Weak
Particle Movement	Vibrational	Flow	Free motion
Characteristic	Stability	Fluidity	Diffusion
Examples	Iron	Water	Nitrogen



Difference Between Gas and Vapor

- **Gas:** Exists in the gaseous state at standard room temperature (25°C).
- **Vapor:** Exists in the gaseous state at a higher temperature, such as water vapor at 100°C.

Matter and Its Properties

The Properties of Matter

Matter can be characterized by its **physical** or **chemical properties**.

1. Physical Properties

- These are the attributes of matter that can be measured or observed without changing its composition or identity.

Examples of Physical Properties:

Color

Shape

Length

Volume

Mass

Density

Classification of Physical Properties

1. Extensive Properties (Quantitative Properties):

1. Depend on the amount of matter present.
2. Examples: Length, Volume, Mass.

2. Intensive Properties (Qualitative Properties):

1. Do not depend on the amount of matter present.
2. Examples: Color, Shape, Density, Boiling Point.

Matter and Its Properties

The Properties of Matter

2. Chemical Properties

- Chemical properties describe the characteristics of a substance when it undergoes a chemical reaction or transformation.

Examples of Chemical Properties:

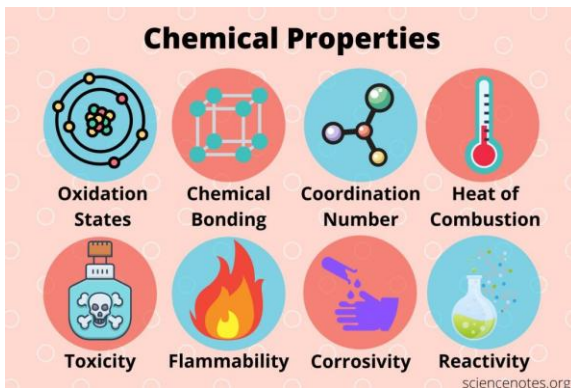
Chemical Reactivity

Oxidation

Combustion

Digestion

Electrolysis



Physical and Chemical Properties of Copper (Cu)

Physical Properties

Reddish-brown color, conductive to heat and electricity

Density: 8.29 g/cm^3

Melting Point: 1085°C

Chemical Properties

Reacts with ammonia to form a deep blue solution

Reacts with moist air to form green copper carbonate compounds

Reacts with nitric acid to form copper nitrate with a blue color

Matter and Its Properties

❑ The Changes of Matter

There are two types of changes that occur to matter: **physical changes** and **chemical changes**

1. Physical Change

A physical change occurs when the appearance or state of the matter changes without altering its chemical composition.

Characteristics of Physical Changes:

- ❖ Physical changes do not produce new substances but retain the original substance.
- ❖ All changes in the states of matter are classified as physical changes.

CHANGING STATES OF MATTER



Examples of State Changes (Physical Changes):

Change Type	Definition
Melting	Conversion from solid to liquid.
Evaporation	Conversion from liquid to gas.
Freezing	Conversion from liquid to solid.
Condensation	Conversion from gas to liquid.
Sublimation	Conversion directly from solid to gas.

Matter and Its Properties

❑ The Changes of Matter

There are two types of changes that occur to matter: **physical changes** and **chemical changes**

2. Chemical Change (Chemical Reaction)

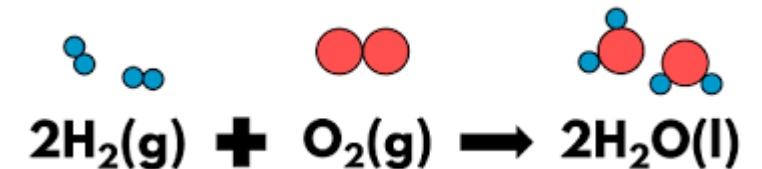
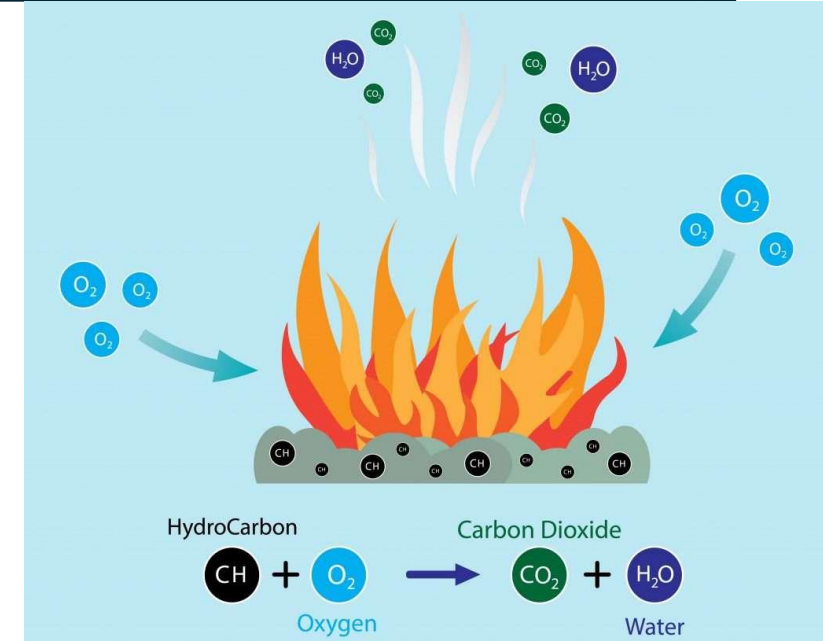
A chemical change involves an alteration in the internal composition of matter, resulting in the transformation of the substance into a different one.

Characteristics of a Chemical Change:

- ❖ A chemical change produces new substances with properties that differ completely from the original substance.
- ❖ Examples include various types of chemical reactions such as combustion, oxidation, and decomposition reactions.

Indicators of a Chemical Reaction:

1. Gas evolution
2. Precipitate formation
3. Color change
4. Temperature change
5. Light emission or sparks during combustion

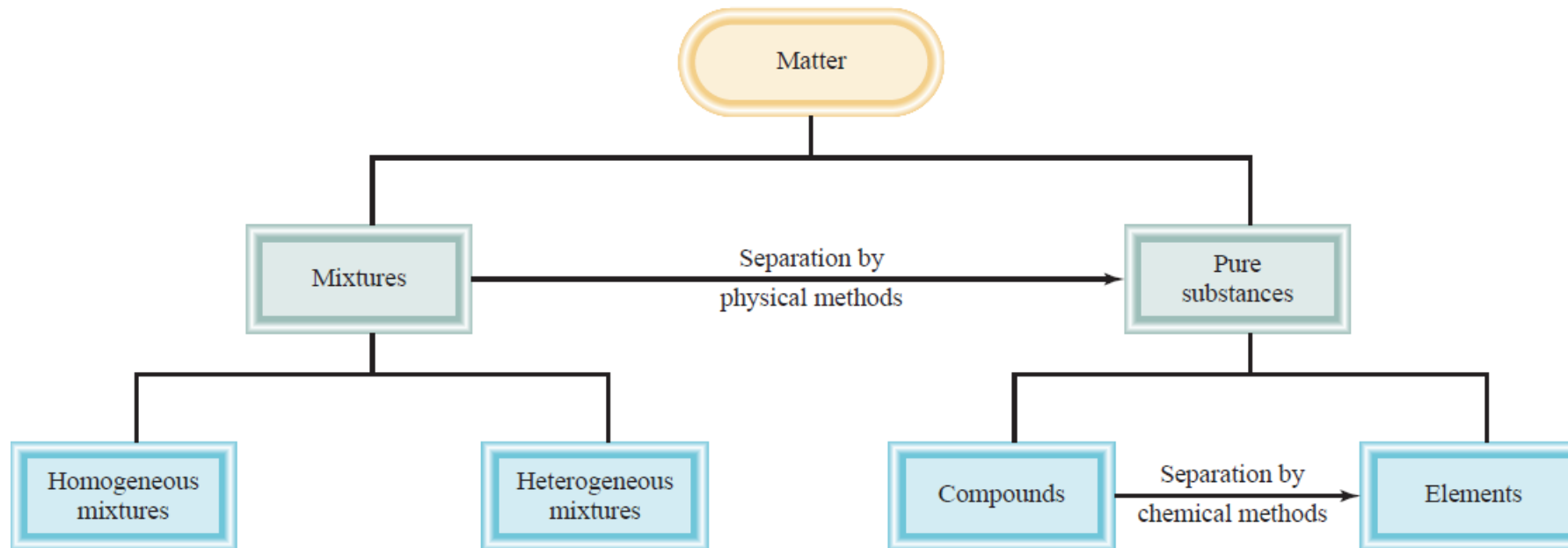


Examples of chemical changes such as methane combustion and water formation.

Matter and Its Properties

❑ The classification of Matter

Matter exists in two main forms: **pure substances** and **impure substances (mixtures)**.



Pure Substances:

- Composed of only one type of component and has a fixed composition.
- Cannot be separated into simpler substances by physical means.

Impure Substances (Mixtures):

- Composed of two or more substances that retain their individual properties.
- Can be separated by physical methods.

Matter and Its Properties

Pure Substances

1. **Elements:** An element is a basic substance that cannot be broken down into simpler substances by chemical means.

➤ Scientists have discovered 114 elements.

Names and Symbols of Some Elements

•The English or Latin names of elements are used to name them.

•The scientist Berzelius was the first to use symbols to represent elements, choosing one or two letters from the English or Latin name of the element according to specific rules:

- If the element's symbol consists of one letter, it is written as a **capital letter**.
- If the element's symbol consists of two letters, the first is a **capital letter**, and the second is a **small letter**.
- In some cases, the Latin name is used for the element's symbol.

Some Common Elements and Their Symbols

Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bromine	Br	Iodine	I	Potassium	K
Calcium	Ca	Iron	Fe	Silicon	Si
Carbon	C	Lead	Pb	Silver	Ag
Chlorine	Cl	Magnesium	Mg	Sodium	Na
Chromium	Cr	Mercury	Hg	Sulfur	S
Cobalt	Co	Nickel	Ni	Tin	Sn
Copper	Cu	Nitrogen	N	Zinc	Zn

Matter and Its Properties

Pure Substances

2. Compounds

A compound is a substance formed from the chemical combination of two or more elements in fixed-weight ratios.

- **Example:** The reaction of sodium (a highly reactive metal) with chlorine gas (a toxic gas) results in the formation of sodium chloride (table salt), which is a stable compound.

Properties of Compounds:

1. Formed as a result of a chemical reaction.
2. The properties of the compound differ entirely from those of its components.
3. Composed of multiple elements in fixed weight ratios.
4. Can only be broken down into its components by chemical methods.



Law of Definite Proportions (Constant Proportions Law):

Every chemical compound, regardless of the method of preparation, is composed of the same elements in a fixed ratio by weight.

Formula:

$$\text{Percentage of Element} = \frac{\text{Mass of Element}}{\text{Mass of Compound}} \times 100$$

Matter and Its Properties

Mixtures

Definition of Mixture:

A mixture is a combination of two or more substances that are mixed together without undergoing a chemical reaction.

• Mixtures are classified into two types:

- **Homogeneous Mixtures:**

- The components are distributed uniformly and cannot be distinguished by the naked eye.
- Examples: True solutions, colloidal solutions.

- **Heterogeneous Mixtures:**

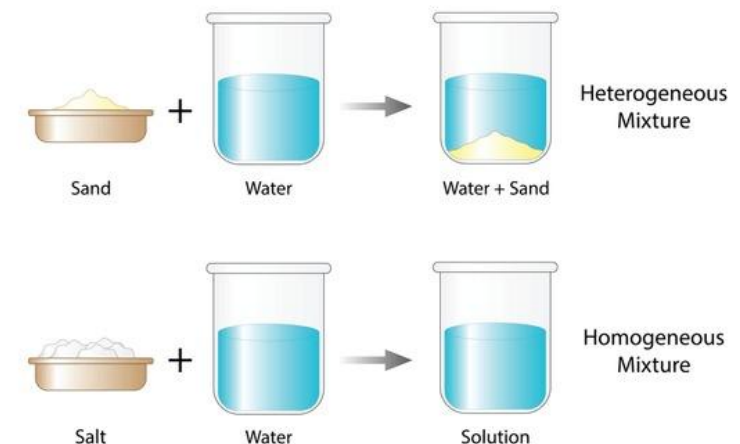
- The components are not uniformly distributed and can be distinguished by the naked eye.
- Examples: Suspensions.

Properties of Mixtures:

- No chemical reaction occurs between the components.
- Each substance retains its original properties.
- The substances can be combined in any ratio.
- Mixtures can be separated into their components using physical methods.

Components of Solutions:

- **Solvent:** Present in a larger proportion.
- **Solute:** Present in a smaller proportion.



Matter and Its Properties

Methods of Separating Matter

Two types of methods are used to separate matter: **physical** and **chemical**.

1. Physical Methods:

Physical methods involve the separation of components without changing their chemical composition.

Examples of Physical Separation Methods

Types of mixtures	Mixtures	Methods of separation
Heterogeneous	Solid and solid	Handpicking, sieving, winnowing, magnetic separation, sublimation.
	Insoluble solid and liquid	Sedimentation and decantation, loading, filtration, centrifugation
	Two immiscible liquids	Decantation, solvent extraction
Homogeneous	Soluble solid and liquids	Evaporation, distillation, crystallisation
	Two miscible liquids	Fractional distillation
	Solution of two or more solids in a liquid	Chromatography

Matter and Its Properties

Methods of Separating Matter

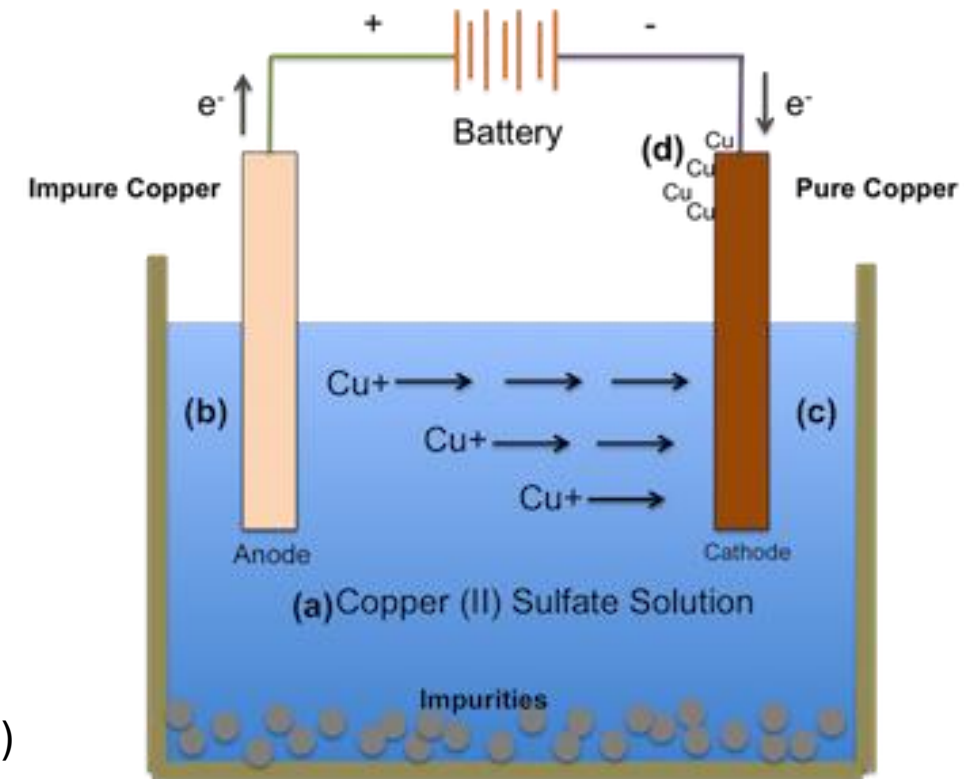
Two types of methods are used to separate matter: **physical** and **chemical**.

2. Chemical Methods (Electrolysis Process):

Chemical methods involve a change in the composition of the substance during the separation process. One example is **electrolysis**.

Electrolysis Process:

- ☐ Some compounds can be broken down into their elements by the flow of an electric current. This process is called **electrolysis**.
- ☐ Water can be decomposed into its components (oxygen and hydrogen gases) through electrolysis.
- ☐ The process does not occur spontaneously; it requires external energy to drive the reaction.
- ☐ An **electrolytic cell** consists of two electrodes (anode and cathode) immersed in an electrolyte (e.g., water containing ions). This process is known scientifically as **Hofmann's voltameter**.
- ☐ The gases produced (hydrogen and oxygen) are in a **fixed ratio of 2:1**, consistent with the molecular formula of water (H_2O).



Measurement and Handling the Numbers

Contents

Contents

•Introduction

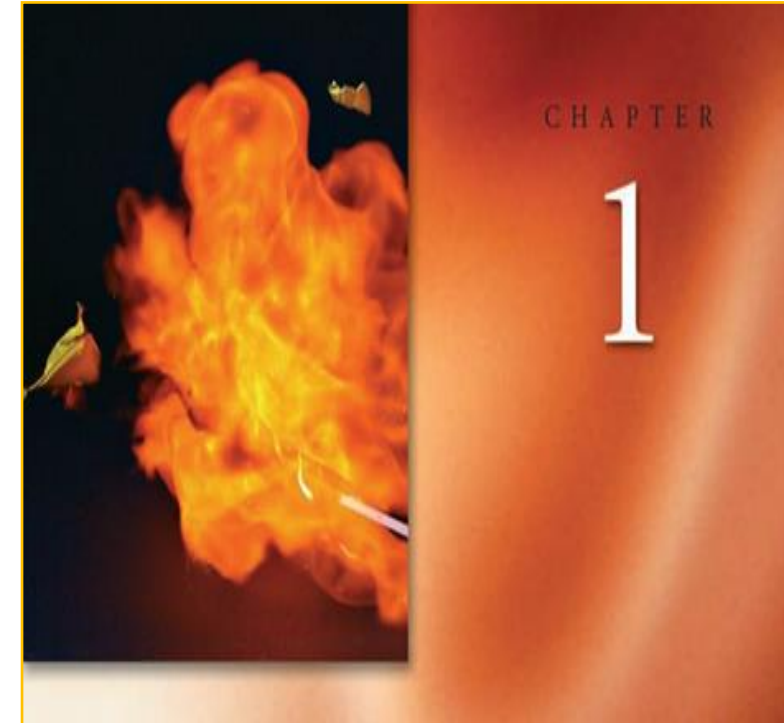
- Definition of the measurement process
- Elements of the measurement process
- How to represent measured quantities

•Numbers

- Scientific notation for numbers
- Measurement constants and their significance
- Significant figures and their notation

•Units of Measurement

- The International System of Units (SI)
- Basic measurement units
- Some derived measurement units



Measurement

Definition of Measurement

It is the comparison of an unknown quantity with a known (standard) quantity.

The measurement process consists of three main elements:

Measured Quantities – Units of Measurement – Measuring tools

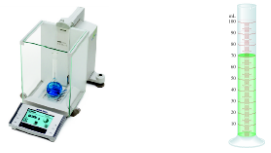
➤ Measured Quantities

➤ Fundamental Quantities

Quantities that cannot be derived from other quantities.
e.g., length, mass, time, etc.

➤ Derived Quantities:

Quantities that can be derived from other fundamental quantities.
e.g., volume, density, pressure.



Units of Measurement

➤ Fundamental Units:

The units used to express fundamental quantities.

➤ Derived Units:

The units used to express derived quantities.

Measuring Instruments (tools)

Devices, whether electronic or manual, used to measure quantities as a specific standard or scale.

Examples of Fundamental and Derived Quantities

Table 1.2 SI Base Units

Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

Notes

- ✓ There are many measurement systems, such as the British and French systems.
- ✓ The **International System of Units (SI Units)** is a globally accepted and standardized system used by scientists worldwide.

Measurement

Representation of Measured Quantities (Numbers)

Any measurable quantity can be represented by a **number** and a **measurement unit**.

Number

Represents the measured quantity. It is meaningless without a measurement unit.

Measurement Unit (Synthetic Matter)

Used to distinguish the written number.

Example:

•Length of :

53.3 cm →

- **Number:** 53.3
- **Unit:** cm

•Mass (Weight):

3.9 kg →

- **Number:** 3.9
- **Unit:** kg

Representation of Measured Quantities (Numbers)

First: Scientific Notation for Numbers

- ❖ Scientific notation is used to handle very large or very small numbers because dealing with these numbers directly can lead to errors. These numbers are represented using the following formula:

$$N \times 10^n$$

Where:

N: A number between 1 and 10

n: An integer (positive or negative)

- If we have a number and are asked to write it in scientific notation, we follow these steps:

1. Determine the value of the exponent (**n**) by identifying the number of decimal places the point needs to be moved to give **N** (a number between 1 and 10):

1. If the decimal point is moved **to the left**, **n** is positive.
2. If the decimal point is moved **to the right**, **n** is negative.

Example

Write the following numbers in scientific notation:

Number	Scientific Notation
22.2233	2.22233×10^1
0.003	3.0×10^{-3}
15534.66	1.553466×10^4
0.00000772	7.72×10^{-6}
568.76	5.6876×10^2

Representation of Measured Quantities (Numbers)

1. Scientific Notation for Numbers

Arithmetic Operations in Scientific Notation

Addition

Write each quantity with the same exponent, then add the bases:

$$(N1 \times 10^n) + (N2 \times 10^n) = (N1 + N2) \times 10^n$$

Subtraction

Write each quantity with the same exponent, then subtract the bases:

$$(N1 \times 10^n) - (N2 \times 10^n) = (N1 - N2) \times 10^n$$

Multiplication

Multiply the bases and add the exponents:

$$(N1 \times 10^n) \times (N2 \times 10^m) = (N1 \times N2) \times 10^{n+m}$$

Division

Divide the bases and subtract the exponents:

$$(N1 \times 10^n) / (N2 \times 10^m) = (N1 / N2) \times 10^{n-m}$$

Operation	Calculation	Result
Subtraction	$(2.22 \times 10^{-2}) - (4.10 \times 10^{-3})$	1.81×10^{-2}
Addition	$(4.31 \times 10^4) + (3.9 \times 10^3)$	4.70×10^4
Addition	$(7.4 \times 10^3) + (2.1 \times 10^3)$	9.5×10^3
Division	$(6.9 \times 10^7) / (3.0 \times 10^{-5})$	2.3×10^{12}
Division	$(8.5 \times 10^4) / (5.0 \times 10^9)$	1.7×10^{-5}
Multiplication	$(8.0 \times 10^4) \times (5.0 \times 10^2)$	4.0×10^7

Representation of Measured Quantities (Numbers)

2. Prefixes Used in Measurement and Their Denotations

Prefixes are segments added at the beginning of a unit name to facilitate conversion between units in the International System of Units (SI).

TABLE 1.3 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
tera-	T	1,000,000,000,000, or 10^{12}	1 terameter (Tm) = 1×10^{12} m
giga-	G	1,000,000,000, or 10^9	1 gigameter (Gm) = 1×10^9 m
mega-	M	1,000,000, or 10^6	1 megameter (Mm) = 1×10^6 m
kilo-	k	1,000, or 10^3	1 kilometer (km) = 1×10^3 m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	μ	1/1,000,000, or 10^{-6}	1 micrometer (μ m) = 1×10^{-6} m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) = 1×10^{-9} m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) = 1×10^{-12} m

Representation of Measured Quantities (Numbers)

3. Significant Figures

- Significant figures represent the digits in a measured or calculated quantity that reflect its precision.
- The determination of significant figures refers to the percentage of error in measurement or calculation.
 - ✓ **The more significant figures, the higher the precision of the measurement.**

Rules for Determining Significant Figures

Rule	Examples	Number of Significant Figures
1. Any digit that is not zero is significant.	845 cm	3
2. Zeros between nonzero digits are significant	606 m	3
	40501 kg	5
3. If a number is greater than 1, then all zeros to the right of the decimal point are significant.	2.0 mg	2
	40.062 mL	5
4. If a number is less than 1, then only the zeros that are at the end and in the middle of the number are significant.	0.00420 g	3
• Zeros to the left of the first nonzero digit are not significant	0.08 L	1

How many significant figures are in each of the following measurements?

24 mL

3001 g

0.0320 m³

6.4×10^4 molecules

560 kg

How many significant figures are in each of the following measurements?

24 mL	2 significant figures
3001 g	4 significant figures
0.0320 m ³	3 significant figures
6.4×10^4 molecules	2 significant figures
560 kg	2 significant figures

Representation of Measured Quantities (Unit of Measurement)

1. Measuring Length

Length is measured using tools like the **ruler** or **measuring tape** as basic tool.

Units of Length Measurement

Meter	Symbol	Conversion Between Length Units
Kilometer	km	1 km = 1000 m
Centimeter	cm	1 m = 100 cm = 1000 mm
Millimeter	mm	1 m = 1×10^3 mm
Nanometer	nm	1 m = 1×10^9 nm
Picometer	pm	1 m = 1×10^{12} pm



Conversion Notes:

- 1 km = 1000 m
- 1 m = 100 cm
- 1 m = 1×10^9 nm
- 1 m = 1×10^{12} pm

Representation of Measured Quantities (Unit of Measurement)

2. Measuring Mass

•**Definition:** measure of the quantity of matter

➤ Mass is measured using different types of **balances** as basic tools for mass measurement.

Difference Between Mass and Weight:

•**Mass:** A measure of the amount of matter in an object.

•**Weight:** The force exerted by gravity on an object.

➤ Mass does not depend on the location of the object.

➤ Weight depends on the gravitational force at the object's location.

Formula for Weight:



Fig. A sensitive balance used to measure mass

SI unit of mass is the **kilogram** (kg)

$$1 \text{ kg} = 1000 \text{ g} = 1 \times 10^3 \text{ g}$$

Units of Mass Measurement

Unit	Symbol	Conversion Between Mass Units
Kilogram	kg	
Gram	g	1 kg = 1000 g
Milligram	mg	1 g = 1000 mg
Microgram	μg	1 g = 1,000,000 μg

Representation of Measured Quantities (Unit of Measurement)

3. Measuring Time

The **clock** is used as a basic tool for measuring time.



Notable Achievement:
Ahmed Zewail, a scientist, won the **1999 Nobel Prize in Chemistry** for his invention of the **femtosecond laser microscope**, which captures chemical reactions in incredibly short time intervals.



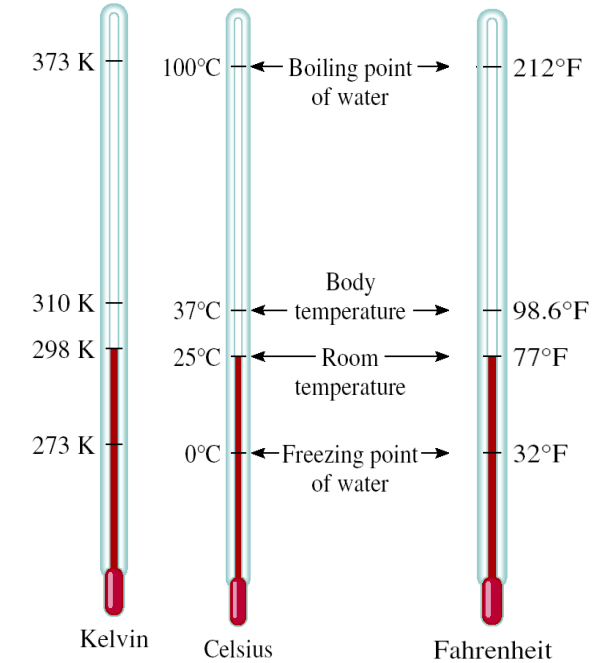
Units of Time Measurement

Unit	Symbol	Conversion Between Time Units
Hour	h	1 h = 60 min
Minute	min	1 min = 60 s
Nanosecond	ns	1 s = 1×10^9 ns
Femtosecond	fs	1 s = 1×10^{15} fs

Representation of Measured Quantities (Unit of Measurement)

4. Measuring Temperature

The **thermometer** is used as a basic tool for measuring temperature.



Units of Temperature Measurement

Unit	Symbol	Conversion Between Temperature Units
Kelvin	K	$K = ^\circ C + 273.15$
Celsius	$^\circ C$	$^\circ C = K - 273.15$
Fahrenheit	$^\circ F$	$^\circ F = 1.8 \times ^\circ C + 32$
		$^\circ C = (^\circ F - 32) / 1.8$

- Boiling point of water:
100°C, 373 K, 212°F
- Freezing point of water:
0°C, 273 K, 32°F
- Absolute zero:
0 K, -273°C, -459.67°F

$$273 \text{ K} = 0 \text{ }^\circ\text{C}$$
$$373 \text{ K} = 100 \text{ }^\circ\text{C}$$

$$K = ^\circ C + 273.15$$
$$^\circ F = \frac{9}{5} \times ^\circ C + 32$$
$$32 \text{ }^\circ\text{F} = 0 \text{ }^\circ\text{C}$$
$$212 \text{ }^\circ\text{F} = 100 \text{ }^\circ\text{C}$$

Representation of Measured Quantities (Unit of Measurement)

Problems on Conversion Between Temperature Units

Example 1: Mercury is the only metal that exists in liquid form at room temperature. It melts at -38.9°C . Convert the melting point to Kelvin (K).

Example 2:

Solder (a mixture of tin and lead) melts at 224°C . Convert the melting point to Fahrenheit.

Example 3:

Helium has the lowest boiling point among all elements in the periodic table, at -245°F . Convert this temperature to Celsius

Representation of Measured Quantities (Unit of Measurement)

Conversion Between Temperature Units

Example 1: Mercury is the only metal that exists in liquid form at room temperature. It melts at -38.9°C . Convert the melting point to Kelvin (K).

Solution:

$$K = (^{\circ}\text{C} + 273.15)$$

$$K = (-38.9 + 273.15) = 234.15 \text{ K}$$

Example 2:

Solder (a mixture of tin and lead) melts at 224°C . Convert the melting point to Fahrenheit.

Solution:

$$^{\circ}\text{F} = (1.8 \times ^{\circ}\text{C}) + 32$$

$$^{\circ}\text{F} = (1.8 \times 224) + 32 = 435^{\circ}\text{F}$$

Example 3:

Helium has the lowest boiling point among all elements in the periodic table, at -245°F . Convert this temperature to Celsius

Solution:

$$^{\circ}\text{C} = (^{\circ}\text{F} - 32) \div 1.8$$

$$^{\circ}\text{C} = (-245 - 32) \div 1.8$$

$$^{\circ}\text{C} = -277 \div 1.8 = -153.8^{\circ}\text{C}$$

Representation of Measured Quantities (Unit of Measurement)

5. Measuring the Amount of Substance

➤ The mole is the quantity that contains Avogadro's number of units (atoms, molecules, ions, etc.).

• Equations for Calculating Moles:

- $n = \frac{N}{N_A}$
- n : Number of moles
 - N : Number of chemical units (atoms, molecules, ions)
 - N_A : Avogadro's number (6.0221415×10^{23})

- $n = \frac{m}{M}$
 - n : Number of moles
 - m : Mass of the substance (g)
 - M : Molar mass (g/mol)

Notes:

- Avogadro's number is critical for conversions between number of particles and moles.
- The molar mass is the mass of 1 mole of a substance in grams.

Units of Measurement for the Amount of Substance

Unit	Symbol	Conversions
Mole	mol	1 mol=1000 mmol
Millimole	mmol	
Micromole	μmol	1 mol=1,000,000 μmol

Problems

Calculate the number of moles in 1500 molecules of sucrose ($\text{C}_{11}\text{H}_{22}\text{O}_{11}$).

Solution:

$$n = \frac{N}{N_A} = 1500 / 6.022 \times 10^{23} = 2.49 \times 10^{-21} \text{ mol}$$

How many moles are in 48 g of oxygen atoms (O)?

Solution:

$$n = \frac{m}{M} = 48 / 16 = 3 \text{ mol}$$

How many moles are in 48 g of oxygen molecules (O_2)?

Solution:

$$n = \frac{m}{M} = 48 / 32 = 1.5 \text{ mol}$$

Representation of Measured Quantities (Derived Units)

1. Measuring Volume (Volume)

- **Definition:** The volume of an object is defined as the amount of space it occupies.

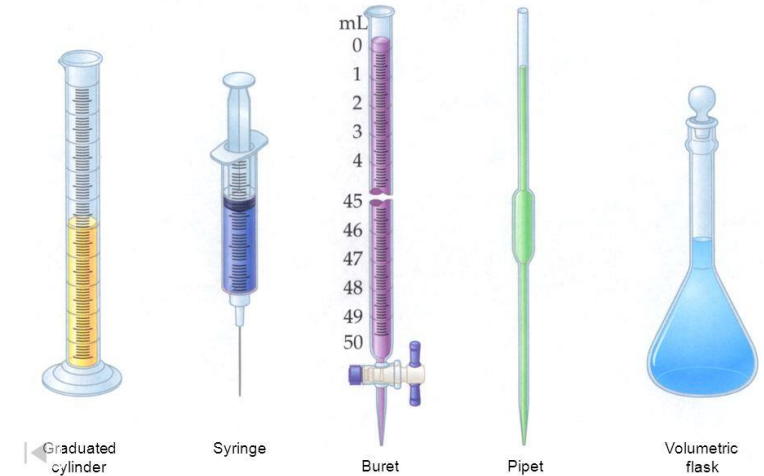
Volume can be calculated or measured using the following relationships:

$$\text{Volume} = L \times W \times H = m \times m \times m = m^3$$

Tools like **graduated cylinders**, **volumetric flasks**, **pipettes**, and **burettes** are used to measure volume in the lab.

The **liter (L)** is commonly used in chemistry for volume measurements and corresponds to the volume of a cubic decimeter (1 dm^3).

Instruments for Measuring Volume



Unit	Symbol	Conversion Between Units
Cubic meter	m^3	
Cubic centimeter	cm^3	$1 \text{ cm}^3 = (1 \times 10^{-2} \text{ m})^3 = 1 \times 10^{-6} \text{ m}^3$
Cubic decimeter	dm^3	$1 \text{ dm}^3 = (1 \times 10^{-1} \text{ m})^3 = 1 \times 10^{-3} \text{ m}^3$
Liter	L	$1 \text{ L} = 1000 \text{ mL} = 1000 \text{ cm}^3 = 1 \text{ dm}^3$
Milliliter	mL	$1 \text{ mL} = 1 \text{ cm}^3$

Representation of Measured Quantities (Derived Units)

2. Measuring Density

Density is defined as the mass of a substance per unit of volume.

Density can be calculated using the relationship:

$$d = \frac{m}{V}, \text{ unit: } \frac{Kg}{m^3}$$

Where: d: Density

- m : Mass
- V : Volume

•The **hydrometer** is commonly used to measure the density of liquids.

Unit	Symbol	Conversion
Kilogram per cubic meter	kg/m ³	1 g/cm ³ = 1 g/mL
Gram per cubic centimeter	g/cm ³	1 kg/m ³ = 0.001 g/mL

Densities of Some Substances at 25°C

Substance	Density (g/cm ³)
Air*	0.001
Ethanol	0.79
Water	1.00
Mercury	13.6
Table salt	2.2
Iron	7.9
Gold	19.3
Osmium [†]	22.6

* Measured at 1 atmosphere.

[†] Osmium (Os) is the densest element known

Problems

Gold is a chemically inert metal used widely in jewelry, electronics, and dentistry. If a gold bar has a mass of 301 g and a volume of 15.6 cm³, calculate its density.

A piece of platinum metal with a density of 21.5 g/cm³ has a volume of 4.49 cm³. What is its mass?

Calculate the mass in kilograms of a solution with a volume of 10 mL and a density of 0.713 g/mL

Problems

Gold is a chemically inert metal used widely in jewelry, electronics, and dentistry. If a gold bar has a mass of 301 g and a volume of 15.6 cm³, calculate its density.

Solution:

$$d = \frac{m}{V} = \frac{301g}{15.6 \text{ cm}^3} = 19.3 \text{ g/cm}^3$$

A piece of platinum metal with a density of 21.5 g/cm³ has a volume of 4.49 cm³. What is its mass?

Solution:

$$m = d \times V = 21.5 \text{ g/cm}^3 \times 4.49 \text{ cm}^3 = 96.53 \text{ g}$$

Calculate the mass in kilograms of a solution with a volume of 10 mL and a density of 0.713 g/mL

Solution:

$$m = d \times V = 0.713 \text{ g/mL} \times 10 \text{ mL} = 7.13 \text{ g} = 0.00713 \text{ kg}$$

Representation of Measured Quantities (Derived Units)

3. Measuring Speed, Acceleration, and Force

Speed is defined as the ratio of distance to time.

$$v = \frac{d}{t} = \frac{m}{s}$$

Acceleration is defined as the rate of change of speed over time.

$$a = \frac{v}{t} = \frac{m}{s \cdot s} = \frac{m}{s^2}$$

Force is the product of mass and acceleration.

$$F = m \cdot a = m \cdot \frac{v}{t} = \text{Kg} \cdot \frac{m}{s^2}$$

Newton (N): The force that acts on a 1 kg object to produce an acceleration of 1 m/s²

Quantity	Unit	Symbol
Speed	Meter per second	m/s
Acceleration	Meter per second squared	m/s ²
Force	Kilogram meter per second squared (Newton)	kg·m/s ²

Representation of Measured Quantities (Derived Units)

4. Measuring Pressure

Pressure is defined as the **force exerted per unit area**.

The formula for pressure is: $P = \frac{F(\text{Force})}{A(\text{Area})}$

P: Pressure
F: Force (measured in Newtons, N)
A: Area (measured in square meters, m²)

The unit of pressure in the SI system is **Pascal (Pa)**: $P = \frac{\text{N}}{\text{m}^2} = \frac{\text{kg}}{\text{m} \cdot \text{s}^2}$

Unit	Symbol	Conversion
Pascal	Pa	1 Pa = 1 N/m ²
Atmosphere	atm	1 atm = 760 mmHg = 760 Torr
Millimeter Mercury	mmHg	
Torr	Torr	1 atm = 101325 Pa
Bar	bar	1 bar = 100000 Pa

Representation of Measured Quantities (Derived Units)

5. Measuring Energy (Energy)

• **Energy** is the capacity to do work.

➤ Energy exists in different forms, including:

- ❖ Kinetic Energy
- ❖ Potential Energy
- ❖ Chemical Energy

$$1 \text{ KJ} = 1000 \text{ J}$$

$$1 \text{ Cal} = 4.18 \text{ J}$$

$$1 \text{ Kcal} = 1000 \text{ Cal}$$

$$1 \text{ eV} = 1.604 \times 10^{-19} \text{ J}$$

➤ **Chemical Reactions** are categorized based on energy as:

- **Exothermic:** Reactions that release energy (e.g., heat).
- **Endothermic:** Reactions that absorb energy.

Problem: An exothermic reaction releases 85.5 KJ of heat. Convert the energy into Joules, Calories, and Electronvolts.

Solution:

1- convert to Joules: $85,500 \text{ J}$

2- convert to Calories: $20,434.99 \text{ Cal}$

3- convert to Electronvolt: $5.3365 \times 10^{28} \text{ eV}$

CHAPTER

2

Atoms, Molecules, and Ions

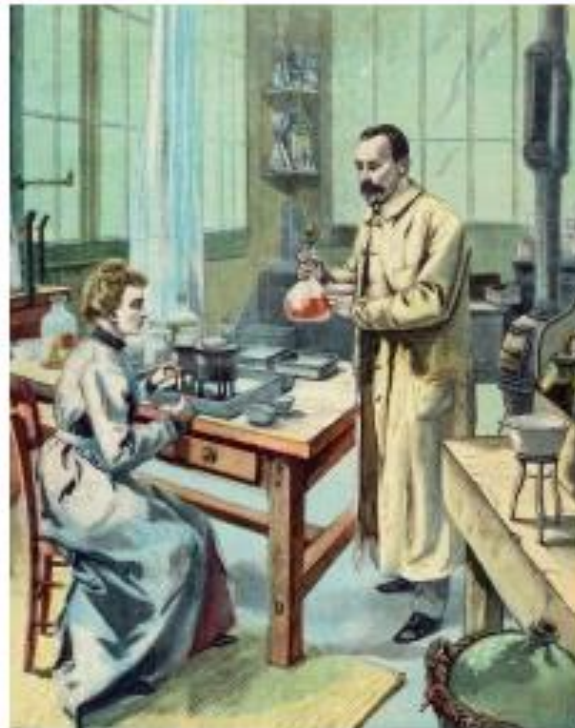


Illustration depicting Marie and Pierre Curie at work in their laboratory. The Curies studied and identified many radioactive elements.

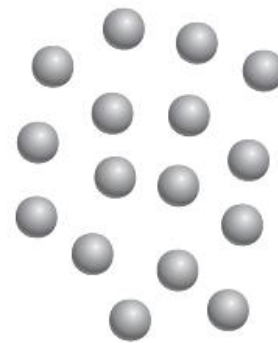
By
Dr. Khalil Mudarmah

The Structure of the Atom

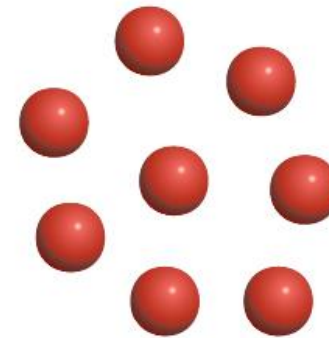
Dalton's atomic theory

1. Elements are composed of extremely small particles called **atoms**.
2. All **atoms** of a given element are identical, having the same size, mass and chemical properties. The atoms of one element are different from the atoms of all other elements.
3. **Compounds** are composed of atoms of more than one element. In any compound, the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.
4. **A chemical reaction** involves only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction.

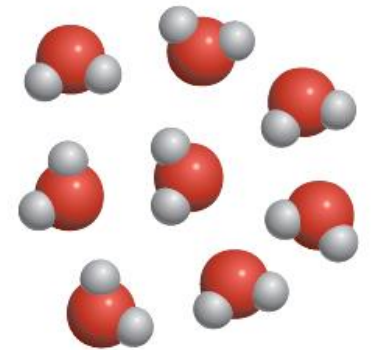
- (a) According to Dalton's atomic theory, atoms of the same element are identical.
- (b) Atoms of one element are different from atoms of other elements. (b) Compound formed from atoms of elements X and Y. In this case, the ratio of the atoms of element X to the atoms of element Y is 2:1. Note that a chemical reaction results only in the rearrangement of atoms, not in their destruction or creation.



Atoms of element X



Atoms of element Y



Compounds of elements X and Y

(a)

(b)

The Structure of the Atom

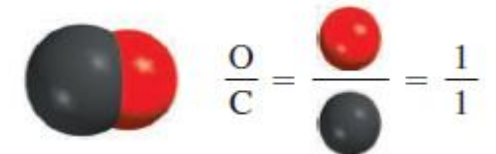
Proust's *law of definite proportions*

“different samples of the same compound always contain its constituent elements in the same proportion by mass.”

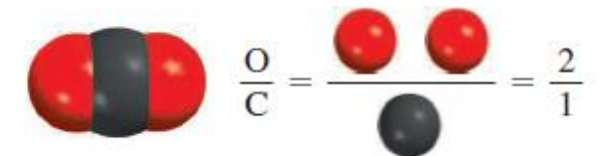
Law of multiple proportions.

According to the law, if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.

Carbon monoxide



Carbon dioxide



Law of Multiple Proportions

The Structure of the Atom

Thomson and the Discovery of Electron

•Cathode Ray Tube:

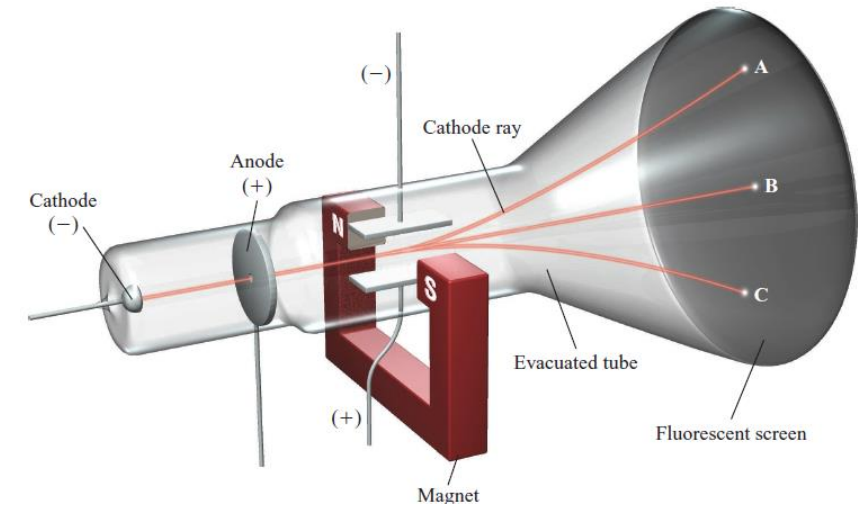
- A glass tube with most air removed, containing two metal plates (cathode and anode) connected to a high-voltage source.
- The cathode emits an invisible ray (cathode ray) drawn to the positively charged anode, producing fluorescence when it strikes a specially coated surface.

•Experimental Setup:

- Additional components: electrically charged plates and a magnet were placed outside the tube.

•Observations:

- With only a magnetic field: the ray strikes point **A**.
- With only an electric field: the ray strikes point **C**.
- When both fields are off or on: the ray strikes point **B**.



•Electromagnetic Interaction:

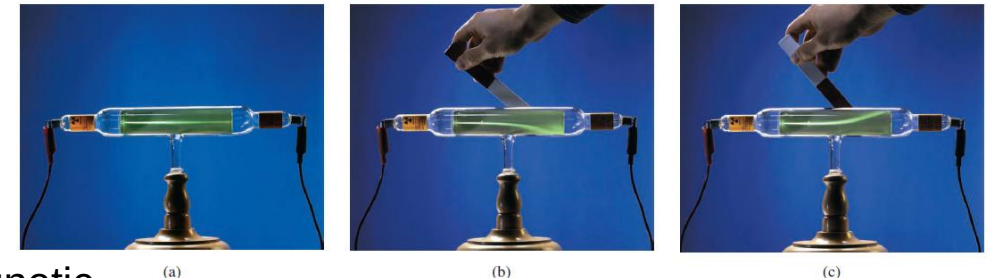
- A moving charged body interacts with electric and magnetic fields.
- Cathode rays are attracted to positively charged plates and repelled by negatively charged ones.

•Discovery of Electrons:

- The behavior of cathode rays confirmed they are composed of negatively charged particles, now known as **electrons**.

•Effect of Magnet:

- A bar magnet deflects the cathode ray, demonstrating its charged and magnetic properties



The Structure of the Atom

Determination the mass of electron

Thomson used a cathode ray tube and his knowledge of electromagnetic theory to determine the ratio of electric charge to the mass of an individual electron. The number he came up with was **$-1.76 \times 10^8 \text{ C/g}$**

where C stands for *coulomb*, which is the unit of electric charge.

Millikan's experiment determined the charge of an electron e^- charge = **$-1.60 \times 10^{-19} \text{ C}$**

$$\begin{aligned}\text{mass of an electron} &= \frac{\text{charge}}{\text{charge/mass}} \\ &= \frac{-1.6022 \times 10^{-19} \text{ C}}{-1.76 \times 10^8 \text{ C/g}} \\ &= 9.10 \times 10^{-28} \text{ g}\end{aligned}$$

The Structure of the Atom

1.Types of rays produced by radioactive decay:

Alpha (α) rays:

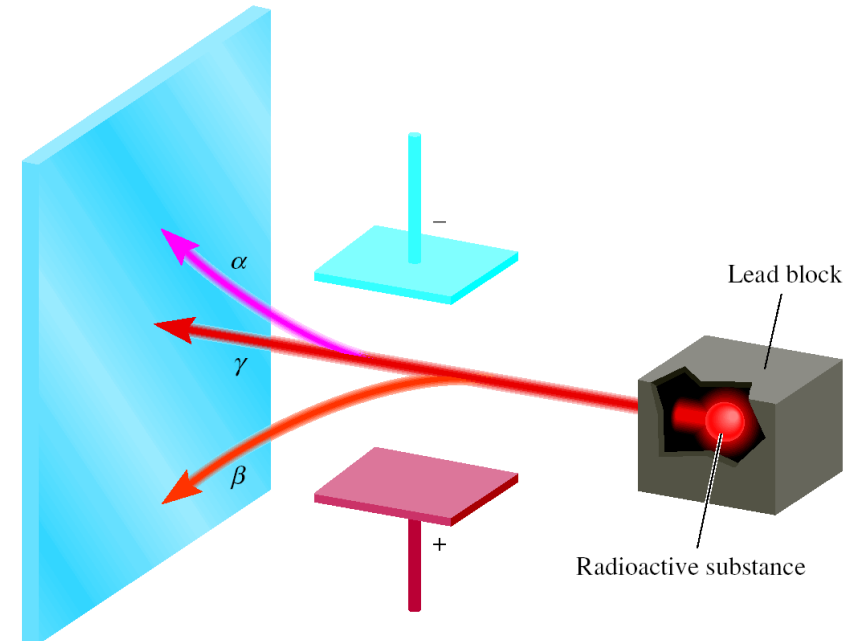
1. Consist of positively charged particles called α particles.
2. Deflected by positively charged metal plates.

Beta (β) rays:

1. Consist of electrons, also known as β particles.
2. Deflected by negatively charged metal plates.

Gamma (γ) rays:

1. High-energy rays similar to X-rays.
2. Have no charge and are not affected by an external field.



Radioactive elements emit three types of rays. rays consist of negatively charged particles (electrons) and are therefore attracted by the positively charged plate. The opposite holds true for α rays— they are positively charged and are drawn to the negatively charged plate. Because gamma Rays have no charges, their path is unaffected by an external electric field.

The Structure of the Atom

Tools Used in Rutherford's Experiment

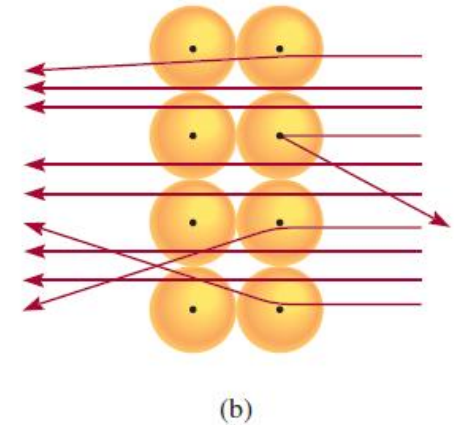
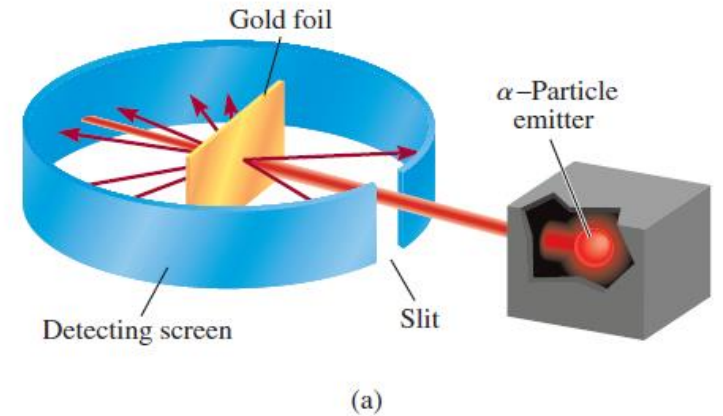
1. **Lead Box:** Contains a source of alpha particles.
2. **Metallic Plate:** Coated with fluorescent material to detect particle collisions (produces flashes of light when hit).
3. **Thin Gold Foil:** Serves as the target for alpha particles.

Steps of the Experiment

1. Alpha particles are directed to collide with the metallic plate.
2. The flashes of light and the number of alpha particles are tracked to determine their deflection patterns.
3. The thin gold foil is placed to expose it to the stream of alpha particles.

Observations

- Most alpha particles pass straight through the gold foil.
- Some alpha particles are deflected at small angles.
- A few alpha particles are deflected at large angles or bounce back, indicating the presence of a dense, positively charged core (the nucleus).



The Structure of the Atom

Rutherford's Model

1.Features of the Atom:

1. Despite its small size, the atom has a complex structure resembling a solar system.
2. It consists of a central nucleus with electrons revolving around it (like planets orbiting the sun).

2.The Nucleus:

1. The nucleus is extremely small, dense, and positively charged.
2. It contains the majority of the atom's mass.

3.Electrons:

1. Electrons are negatively charged and very small in mass compared to the nucleus.
2. They revolve around the nucleus at high speeds in large empty spaces relative to their size.

4.Balance of Charges:

1. The negative charge of the electrons balances the positive charge of the nucleus.
2. This balance ensures the atom is electrically neutral.

5.Electron Motion:

1. Electrons do not collapse into the nucleus due to the balance between centripetal forces and electrostatic forces.

The Structure of the Atom

Limitations of Rutherford's Model:

1.Electron Behavior:

- The model does not explain the system in which electrons revolve around the nucleus or their arrangement.

2.Neutrons:

- Rutherford's model does not mention the presence of neutrons in the atom.

The Structure of the Atom

James Chadwick and the Discovery of Neutrons

- In 1932, James Chadwick was awarded the Nobel Prize for discovering neutrons and determining their charge and mass.

Experiment

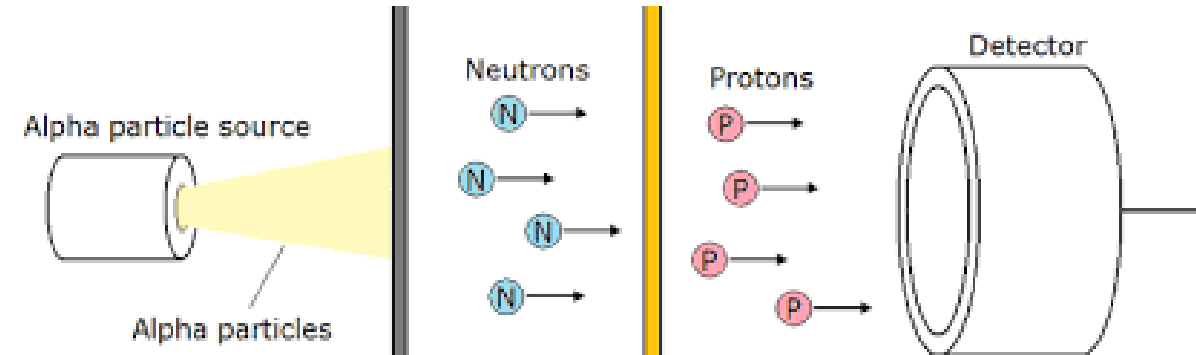
A thin sheet of beryllium was bombarded with alpha (α) particles.

Observation

High-energy rays were emitted from the metal.

Result

These rays were initially termed **gamma rays**, were later identified and named **neutrons**.



Charge and Mass of Neutrons

- When gamma rays passed through a charged plate, they were not deflected, confirming that neutrons are electrically neutral.
- Experiments proved that the mass of a neutron is approximately equal to the mass of a proton.

The Structure of the Atom

The Discovery of Neutrons by James Chadwick

Explanation of the Mass Ratio Between Hydrogen and Helium:

- The nucleus of helium contains **2 protons and 2 neutrons**.
- In contrast, the nucleus of hydrogen contains **only 1 proton and no neutrons**.
- This explains why the mass ratio of helium to hydrogen is **4:1** rather than **2:1**.
- Hydrogen is the only element in the periodic table whose nucleus does not contain neutrons.
- **Note:**
- The mass of an atom depends on the contents of its nucleus (**protons and neutrons**) and not on the electrons.

The Structure of the Atom

Based on Dalton's atomic theory, the atom can be defined as the smallest indivisible part of an element. However, with the advancement of studies, it became clear that the atom is composed of the following:

1.Nucleus:

1. Contains **positively charged protons**.
2. Contains **neutral neutrons** (no charge).

2.Electrons:

1. **Negatively charged electrons** revolve around the nucleus.

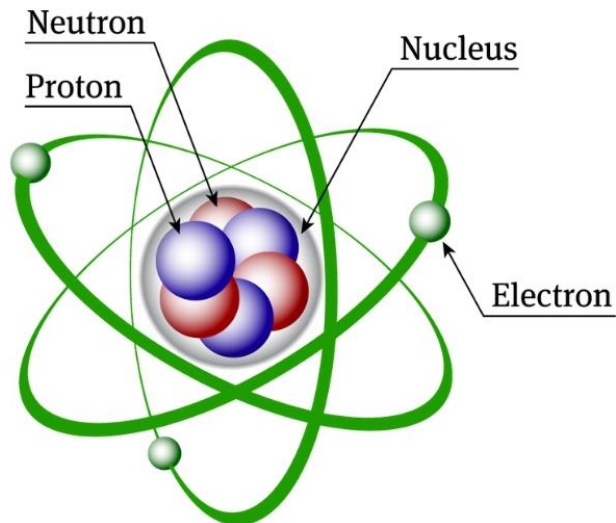
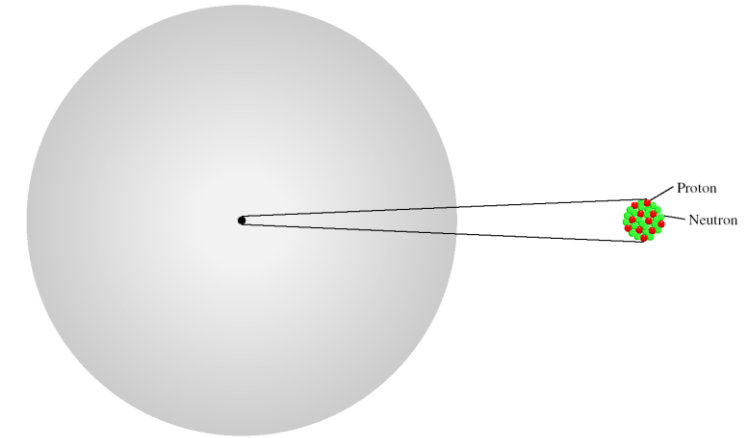


Table 2.1 Mass and Charge of Subatomic Particles

Particle	Mass (g)	Charge	
		Coulomb	Charge Unit
Electron*	9.10938×10^{-28}	-1.6022×10^{-19}	-1
Proton	1.67262×10^{-24}	$+1.6022 \times 10^{-19}$	+1
Neutron	1.67493×10^{-24}	0	0

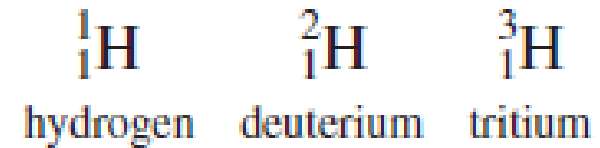
The Structure of the Atom

Atomic Number, Mass Number, and Isotopes

Atomic number (Z) = number of protons in nucleus

Mass number (A) = number of protons + number of neutrons
= atomic number (Z) + number of neutrons

Isotopes are atoms of the same element (X) with different numbers of neutrons in their nuclei



Note:

- Atomic number determines the **chemical identity** of the element.
- All elements (except hydrogen) contain both protons and neutrons in their nuclei.

The Structure of the Atom

How many protons, neutrons, and electrons are in ${}^{14}_6\text{C}$?

How many protons, neutrons, and electrons are in ${}^{11}_6\text{C}$?

Write the chemical symbol for sodium, with Atomic number (Z) = 11, and Mass number (A) = 23.

CHAPTER

2

Atoms, Molecules, and Ions

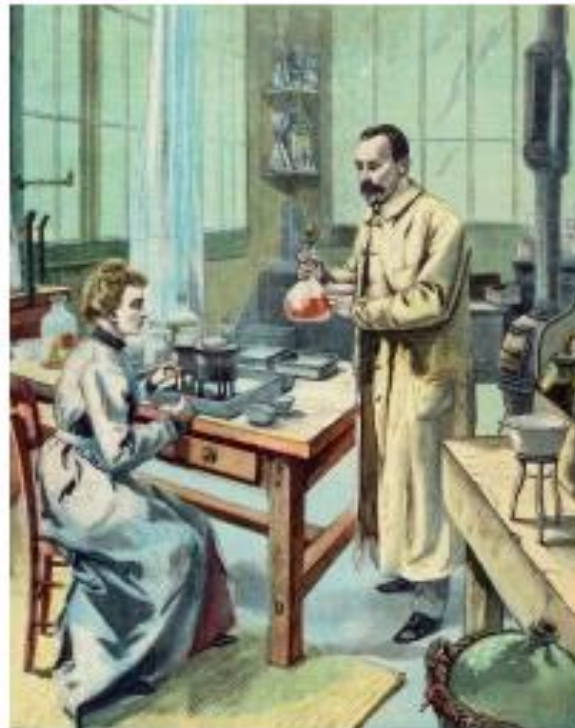


Illustration depicting Marie and Pierre Curie at work in their laboratory. The Curies studied and identified many radioactive elements.

By
Dr. Khalil Mudarmah

The Periodic Table

➤ **Element Discovery (1800-1900):**

- Over half of the known elements were discovered during this period.
- Chemists observed strong similarities among many elements.

➤ **Development of the Periodic Table:**

- Periodic regularities in chemical and physical behavior were recognized.
- The need to organize information on element structures and properties led to the periodic table.
- Elements with similar chemical and physical properties are grouped together.

1 1A																	18 8A
1 H	2 2A											13 3A	14 4A	15 5A	16 6A	17 7A	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113	114 Fl	115	116 Lv	117	118
	Metals																
	Metalloids																
	Nonmetals																
		58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu		
		90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr		

The Periodic Table

➤ **Modern Periodic Table:**

- Arranged by atomic number.
- Organized into:
 - **Horizontal rows:** Called periods.
 - **Vertical columns:** Known as groups or families.
- Elements 113–118 have been synthesized but not named.
- **Elements are categorized into three groups:**

1. Metals - Good conductors of heat and electricity.

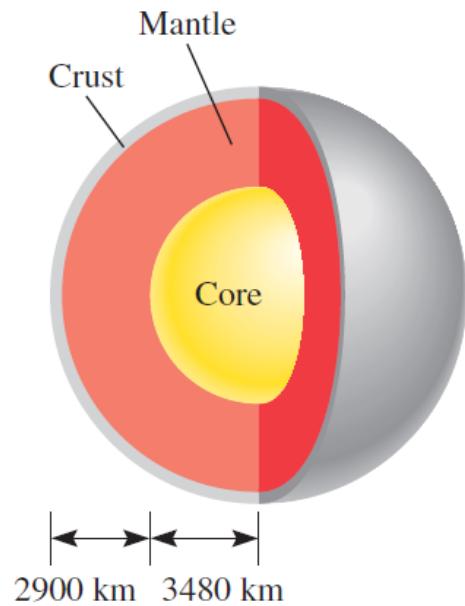
2.Nonmetals - Poor conductors of heat and electricity.

3.Metalloids - Exhibit properties intermediate between metals and nonmetals.

- **Most** elements are metals.
- **17** elements are nonmetals.

1 1A																	18 8A
1 H	2 2A											13 3A	14 4A	15 5A	16 6A	17 7A	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113	114 Fl	115	116 Lv	117	118
	Metals																
	Metalloids																
	Nonmetals																

Distribution of Elements on Earth and in Living Systems

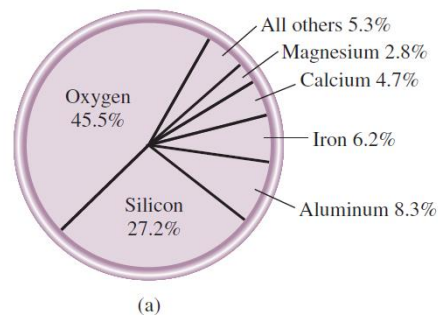


Structure of Earth's interior.

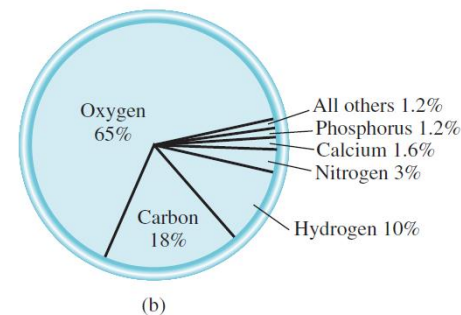
Essential Elements in the Human Body

Element	Percent by Mass*	Element	Percent by Mass*
Oxygen	65	Sodium	0.1
Carbon	18	Magnesium	0.05
Hydrogen	10	Iron	<0.05
Nitrogen	3	Cobalt	<0.05
Calcium	1.6	Copper	<0.05
Phosphorus	1.2	Zinc	<0.05
Potassium	0.2	Iodine	<0.05
Sulfur	0.2	Selenium	<0.01
Chlorine	0.2	Fluorine	<0.01

*Percent by mass gives the mass of the element in grams present in a 100-g sample.



(a) Natural abundance of the elements in percent by mass. For example, oxygen's abundance is 45.5 percent. This means that in a 100-g sample of Earth's crust there are, on the average, 45.5 g of the element oxygen.



(b) Abundance of elements in the human body in percent by mass.

The Periodic Table

Periodic table groups are often referred to by their group numbers (e.g., Group 1A, Group 2A).

➤ **Some groups have special names:**

- **Group 1A:** Alkali metals (Li, Na, K, Rb, Cs, Fr).
- **Group 2A:** Alkaline earth metals (Be, Mg, Ca, Sr, Ba, Ra).
- **Group 7A:** Halogens (F, Cl, Br, I, At).
- **Group 8A:** Noble gases or rare gases (He, Ne, Ar, Kr, Xe, Rn)

The periodic table is organized into columns and rows. The columns are labeled with group numbers and letters: 1A, 2A, 3A, 4A, 5A, 6A, 7A, 8A, 9A, 10A, 11A, 12A, 13A, 14A, 15A, 16A, 17A, 18A. The rows are labeled with period numbers: 1, 2, 3, 4, 5, 6, 7, 8, 9, 10, 11, 12, 13, 14, 15, 16, 17, 18, 19, 20, 21, 22, 23, 24, 25, 26, 27, 28, 29, 30, 31, 32, 33, 34, 35, 36, 37, 38, 39, 40, 41, 42, 43, 44, 45, 46, 47, 48, 49, 50, 51, 52, 53, 54, 55, 56, 57, 58, 59, 60, 61, 62, 63, 64, 65, 66, 67, 68, 69, 70, 71, 72, 73, 74, 75, 76, 77, 78, 79, 80, 81, 82, 83, 84, 85, 86, 87, 88, 89, 90, 91, 92, 93, 94, 95, 96, 97, 98, 99, 100, 101, 102, 103, 104, 105, 106, 107, 108, 109, 110, 111, 112, 113, 114, 115, 116, 117, 118, 119, 120, 121, 122, 123, 124, 125, 126, 127, 128, 129, 130, 131, 132, 133, 134, 135, 136, 137, 138, 139, 140, 141, 142, 143, 144, 145, 146, 147, 148, 149, 150, 151, 152, 153, 154, 155, 156, 157, 158, 159, 160, 161, 162, 163, 164, 165, 166, 167, 168, 169, 170, 171, 172, 173, 174, 175, 176, 177, 178, 179, 180, 181, 182, 183, 184, 185, 186, 187, 188, 189, 190, 191, 192, 193, 194, 195, 196, 197, 198, 199, 200, 201, 202, 203, 204, 205, 206, 207, 208, 209, 210, 211, 212, 213, 214, 215, 216, 217, 218, 219, 220, 221, 222, 223, 224, 225, 226, 227, 228, 229, 230, 231, 232, 233, 234, 235, 236, 237, 238, 239, 240, 241, 242, 243, 244, 245, 246, 247, 248, 249, 250, 251, 252, 253, 254, 255, 256, 257, 258, 259, 260, 261, 262, 263, 264, 265, 266, 267, 268, 269, 270, 271, 272, 273, 274, 275, 276, 277, 278, 279, 280, 281, 282, 283, 284, 285, 286, 287, 288, 289, 290, 291, 292, 293, 294, 295, 296, 297, 298, 299, 300, 301, 302, 303, 304, 305, 306, 307, 308, 309, 310, 311, 312, 313, 314, 315, 316, 317, 318, 319, 320, 321, 322, 323, 324, 325, 326, 327, 328, 329, 330, 331, 332, 333, 334, 335, 336, 337, 338, 339, 340, 341, 342, 343, 344, 345, 346, 347, 348, 349, 350, 351, 352, 353, 354, 355, 356, 357, 358, 359, 360, 361, 362, 363, 364, 365, 366, 367, 368, 369, 370, 371, 372, 373, 374, 375, 376, 377, 378, 379, 380, 381, 382, 383, 384, 385, 386, 387, 388, 389, 390, 391, 392, 393, 394, 395, 396, 397, 398, 399, 400, 401, 402, 403, 404, 405, 406, 407, 408, 409, 410, 411, 412, 413, 414, 415, 416, 417, 418, 419, 420, 421, 422, 423, 424, 425, 426, 427, 428, 429, 430, 431, 432, 433, 434, 435, 436, 437, 438, 439, 440, 441, 442, 443, 444, 445, 446, 447, 448, 449, 450, 451, 452, 453, 454, 455, 456, 457, 458, 459, 460, 461, 462, 463, 464, 465, 466, 467, 468, 469, 470, 471, 472, 473, 474, 475, 476, 477, 478, 479, 480, 481, 482, 483, 484, 485, 486, 487, 488, 489, 490, 491, 492, 493, 494, 495, 496, 497, 498, 499, 500, 501, 502, 503, 504, 505, 506, 507, 508, 509, 510, 511, 512, 513, 514, 515, 516, 517, 518, 519, 520, 521, 522, 523, 524, 525, 526, 527, 528, 529, 530, 531, 532, 533, 534, 535, 536, 537, 538, 539, 540, 541, 542, 543, 544, 545, 546, 547, 548, 549, 550, 551, 552, 553, 554, 555, 556, 557, 558, 559, 560, 561, 562, 563, 564, 565, 566, 567, 568, 569, 570, 571, 572, 573, 574, 575, 576, 577, 578, 579, 580, 581, 582, 583, 584, 585, 586, 587, 588, 589, 590, 591, 592, 593, 594, 595, 596, 597, 598, 599, 600, 601, 602, 603, 604, 605, 606, 607, 608, 609, 610, 611, 612, 613, 614, 615, 616, 617, 618, 619, 620, 621, 622, 623, 624, 625, 626, 627, 628, 629, 630, 631, 632, 633, 634, 635, 636, 637, 638, 639, 640, 641, 642, 643, 644, 645, 646, 647, 648, 649, 650, 651, 652, 653, 654, 655, 656, 657, 658, 659, 660, 661, 662, 663, 664, 665, 666, 667, 668, 669, 670, 671, 672, 673, 674, 675, 676, 677, 678, 679, 680, 681, 682, 683, 684, 685, 686, 687, 688, 689, 690, 691, 692, 693, 694, 695, 696, 697, 698, 699, 700, 701, 702, 703, 704, 705, 706, 707, 708, 709, 710, 711, 712, 713, 714, 715, 716, 717, 718, 719, 720, 721, 722, 723, 724, 725, 726, 727, 728, 729, 730, 731, 732, 733, 734, 735, 736, 737, 738, 739, 740, 741, 742, 743, 744, 745, 746, 747, 748, 749, 750, 751, 752, 753, 754, 755, 756, 757, 758, 759, 760, 761, 762, 763, 764, 765, 766, 767, 768, 769, 770, 771,

Molecules and Ions

- atoms.
-    
- H_2 H_2O NH_3 CH_4

Molecule

- ✓ An aggregate of at least two atoms in a specific arrangement held by chemical bonds.
- ✓ Can consist of atoms of the same element or different elements in fixed ratios.
- ✓ Molecules are **electrically neutral**.

❑ Diatomic molecules:

Contain **two atoms**.

- ✓ Examples of elements: **H₂**, **N₂**, **O₂**, and Group 7A elements (**F₂**, **Cl₂**, **Br₂**, **I₂**).
- ✓ Examples of compounds: **HCl**, **CO**.

❑ Polyatomic molecules:

- Contain **more than two atoms**.

- **Examples:**

- Same element: O_3 (ozone).
- Different elements: H_2O (water), NH_3 (ammonia).

A simplified periodic table with columns labeled 1A, 2A, 3A, 4A, 5A, 6A, 7A, and 8A. The elements are placed in the following positions:

- 1A:** H (Hydrogen)
- 2A:** (Empty)
- 3A:** (Empty)
- 4A:** (Empty)
- 5A:** N (Nitrogen)
- 6A:** O (Oxygen)
- 7A:** F (Fluorine), Cl (Chlorine), Br (Bromine), I (Iodine)
- 8A:** (Empty)

diatomic elements

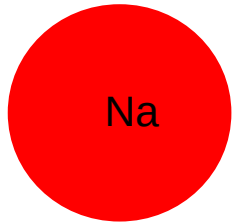
Molecules and Ions

Ion

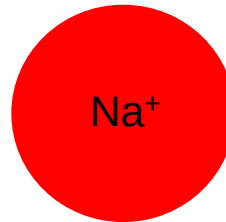
- ✓ An atom or group of atoms with a **net positive or negative charge**.
 - Formed by the **loss or gain of electrons**, while the number of protons in the nucleus remains unchanged.

➤ Cation:

- Positively charged ion.
- Formed when an atom loses one or more electrons.
- Example: Sodium atom (**Na**) loses an electron to form a **sodium cation (Na^+)**.



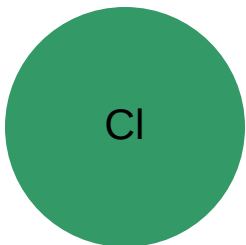
11 protons
11 electrons



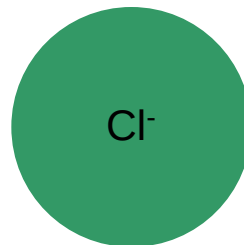
11 protons
10 electrons

➤ Anion:

- Negatively charged ion.
- Formed when an atom gains one or more electrons.
- Example: Chlorine atom (**Cl**) gains an electron to form a **chloride anion (Cl^-)**.



17 protons
17 electrons



17 protons
18 electrons

These processes occur during chemical reactions

Molecules and Ions

Ion

➤ Ionic compounds:

- Formed from **cations** (positive ions) and **anions** (negative ions).
- Example: **Sodium chloride (NaCl)**.

➤ Monatomic ions:

- Contain only **one atom**.
- Examples:
 - **Cations**: Mg^{2+} , Fe^{3+} , Na^+ .
 - **Anions**: S^{2-} , N^{3-} , Cl^- .
- Metals typically form cations, and nonmetals form anions.

➤ Polyatomic ions:

- Contain **two or more atoms** with a net charge.
- Examples:
 - **Hydroxide ion** (OH^-).
 - **Cyanide ion** (CN^-).
 - **Ammonium ion** (NH_4^+).

1 1A	2 2A																18 8A
Li ⁺																	
Na ⁺	Mg ²⁺	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	
K ⁺	Ca ²⁺				Cr ²⁺ Cr ³⁺	Mn ²⁺ Mn ³⁺	Fe ²⁺ Fe ³⁺	Co ²⁺ Co ³⁺	Ni ²⁺ Ni ³⁺	Cu ⁺ Cu ²⁺	Zn ²⁺	Al ³⁺		P ³⁻	S ²⁻	Cl ⁻	
Rb ⁺	Sr ²⁺									Ag ⁺	Cd ²⁺		Sn ²⁺ Sn ⁴⁺		Se ²⁻	Br ⁻	
Cs ⁺	Ba ²⁺									Au ⁺ Au ³⁺	Hg ²⁺ Hg ²⁺		Pb ²⁺ Pb ⁴⁺		Te ²⁻	I ⁻	

Common monatomic ions arranged according to their positions in the periodic table. Note that the Hg_2^{2+} ion contains two atoms

Problems

How many protons and electrons are in ${}_{13}^{27}\text{Al}^{3+}$?

How many protons and electrons are in ${}_{34}^{78}\text{Se}^{2-}$?

Chemical Formulas

❑ Chemical Formulas

✓ Used to represent the composition of molecules and ionic compounds using chemical symbols.

✓ Show the **elements present** and the **ratios** in which atoms combine.

➤ Two types:

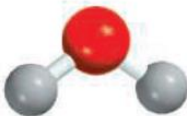
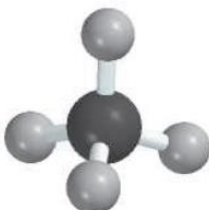
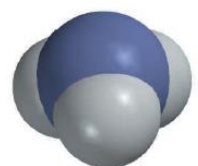
- **Molecular formulas**: Exact number of atoms in a molecule (e.g., H_2 , O_2 , H_2O).
- **Empirical formulas**: Simplest whole-number ratio of elements.

➤ Molecular Formulas

✓ Subscript numerals indicate the number of atoms of each element.

✓ Oxygen (O_2) and ozone (O_3) are **allotropes** of oxygen (distinct forms of the same element).

✓ Carbon has allotropic forms like **diamond** and **graphite**.

	Hydrogen	Water	Ammonia	Methane
Molecular formula	H_2	H_2O	NH_3	CH_4
Structural formula	$\text{H}-\text{H}$	$\text{H}-\text{O}-\text{H}$	$\begin{array}{c} \text{H}-\text{N}-\text{H} \\ \\ \text{H} \end{array}$	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$
Ball-and-stick model				
Space-filling model				

Chemical Formulas

❑ Chemical Formulas

➤ Molecular Models

- Used to visualize molecules since they are too small to observe directly.

➤ Types of models:

- Ball-and-stick models:**

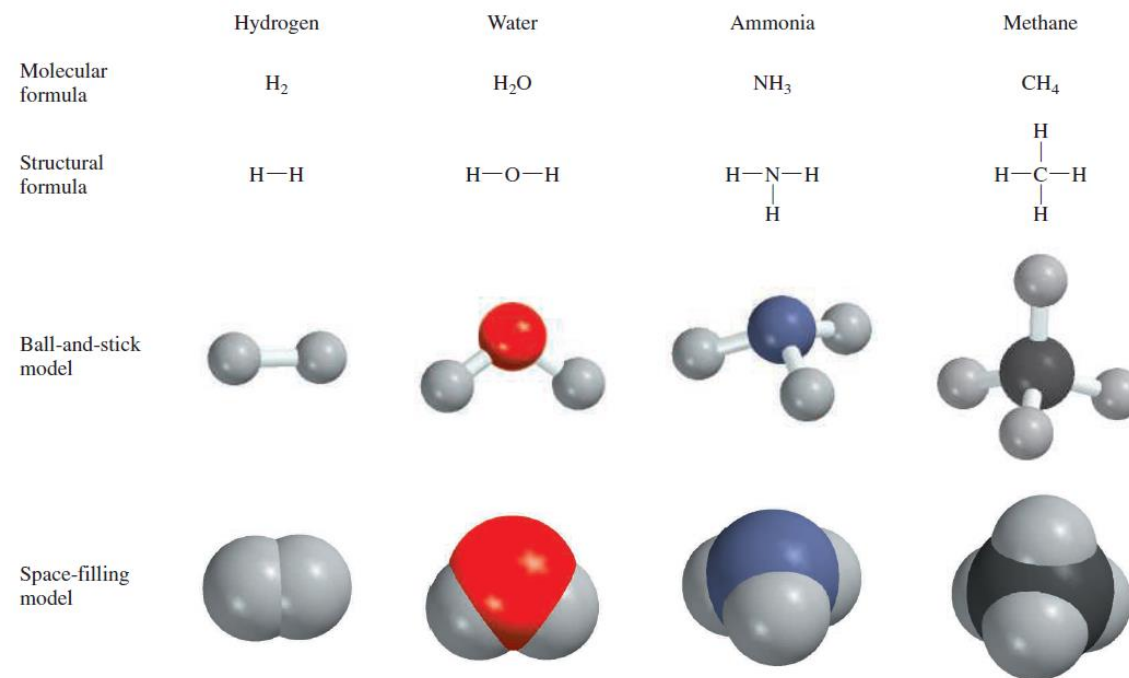
- Atoms are represented by balls, bonds by sticks.
- Shows bond angles and three-dimensional arrangements.
- Limitations: Balls are not proportional to atomic size, and sticks exaggerate space between atoms.

- Space-filling models:**

- Atoms are represented by proportional balls.
- More accurate for atomic size but time-consuming and less clear for 3D arrangement.

➤ Structural Formulas

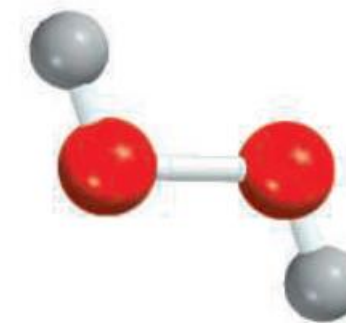
- ✓ Show how atoms are bonded in a molecule (e.g., $\text{H}-\text{O}-\text{H}$ for water).
- ✓ Chemical bonds are represented by lines connecting atomic symbols.



Chemical Formulas

□ Empirical Formula:

- Represents the simplest whole-number ratio of elements in a compound.
- Does not indicate the actual number of atoms in a molecule.
- **Example:**
 - **Hydrogen peroxide:** Molecular formula H_2O_2 , empirical formula HO .
 - **Hydrazine:** Molecular formula N_2H_4 , empirical formula NH_2 .



H_2O_2

❖ Molecular Formula:

- ✓ Represents the actual number of atoms of each element in a molecule.
- ✓ Example: N_2H_4 shows 2 nitrogen and 4 hydrogen atoms.

❖ Relationship Between Molecular and Empirical Formulas:

- Molecular formula gives more detail; the empirical formula is derived by reducing subscripts in the molecular formula.
- Knowing the molecular formula reveals the empirical formula, but not vice versa.

❖ Why Use Empirical Formulas?

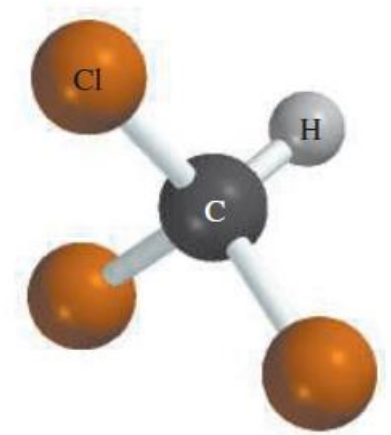
- ✓ Determining the empirical formula is often the first step in analyzing an unknown compound.
- ✓ Additional data are needed to deduce the molecular formula.

• Examples Where Both Formulas Are Identical:

- Water (H_2O), ammonia (NH_3), carbon dioxide (CO_2), methane (CH_4).

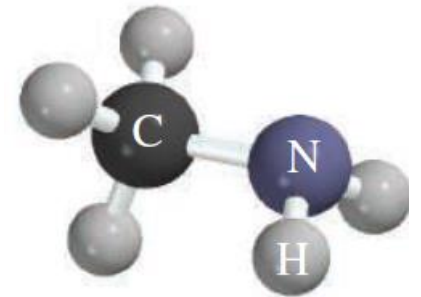
Problems

Write the molecular formula of chloroform, which is used as a solvent and a cleaning agent.



Chloroform

Write the molecular formula of methylamine, a colorless gas used in the production of pharmaceuticals and pesticides, from its ball-and-stick model.



Methylamine

Write the empirical formulas for the following molecules: (a) diborane (B_2H_6), used in rocket propellants; (b) dimethyl fumarate ($\text{C}_8\text{H}_{12}\text{O}_4$), a substance used to treat psoriasis, a skin disease; and (c) vanillin ($\text{C}_8\text{H}_8\text{O}_3$), a flavoring agent used in foods and beverages

Chemical Formulas

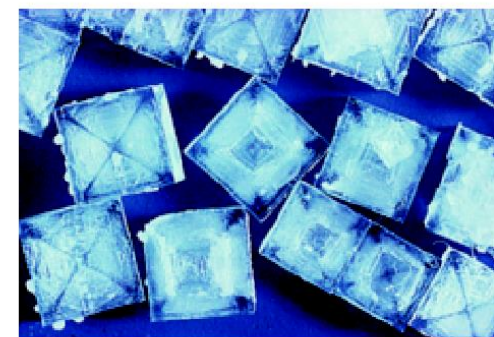
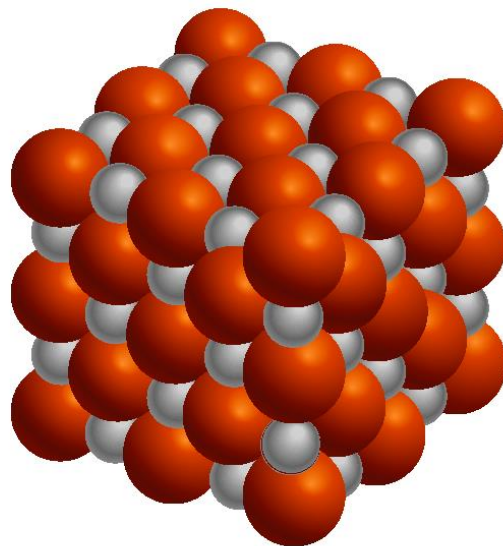
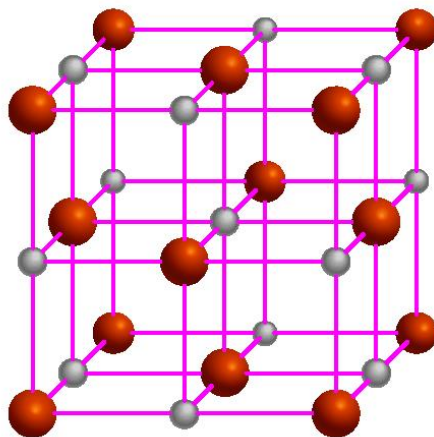
❑ Ionic Compound Formulas

➤ Empirical Formulas:

- ✓ Ionic compounds' formulas are usually their **empirical formulas** because they consist of a 3D network, not discrete molecular units.
- **Example:** Sodium chloride (**NaCl**) has a 1:1 ratio of Na^+ and Cl^- ions.

➤ Electrically Neutral Compounds:

- ✓ The total positive charge (**cations**) and total negative charge (**anions**) in the formula unit must equal **zero**.



The ionic compound NaCl

Chemical Formulas

❑ Ionic Compound Formulas

➤ Rule for Determining Subscripts:

- ✓ The subscript of the **cation** equals the charge on the **anion**.
- ✓ The subscript of the **anion** equals the charge on the **cation**.
- ✓ Subscripts are reduced to their smallest whole-number ratios.

▪ Examples:

❖ Potassium Bromide (KBr):

- ❖ Potassium ion (K^+) and bromide ion (Br^-) have equal charges (+1 and -1).
- ❖ No subscripts are needed; formula = **KBr**.

❖ Zinc Iodide (ZnI_2):

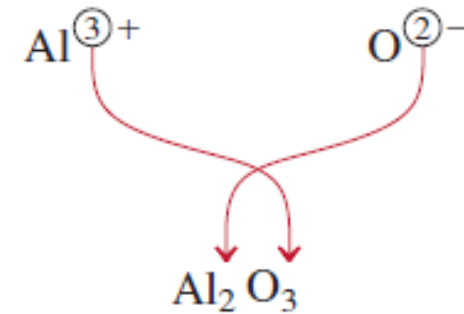
- ❖ Zinc ion (Zn^{2+}) and iodide ion (I^-).
- ❖ Two I^- ions are needed to balance the Zn^{2+} charge; formula = **ZnI₂**.

❖ Aluminum Oxide (Al_2O_3):

- ❖ Aluminum ion (Al^{3+}) and oxide ion (O^{2-}).
- ❖ Two Al^{3+} ions balance three O^{2-} ions; formula = **Al₂O₃**.

1A																		8A
	2A									3A	4A	5A	6A	7A				
Li												N	O	F				
Na	Mg									Al			S	Cl				
K	Ca													Br				
Rb	Sr													I				
Cs	Ba																	

The most reactive **metals** (green) and the most reactive **nonmetals** (blue) combine to form ionic compounds.



Problems

- ❖ Magnesium nitride is used to prepare Borazon, a very hard compound employed in cutting tools and machine parts. Write the formula of magnesium nitride, containing the Mg^{2+} and N^{3-} ions.
- ❖ Write the formulas of the following ionic compounds: (a) chromium sulfate (containing the Cr^{3+} and SO_4^{2-} ions) and (b) titanium oxide (containing the Ti^{4+} and O^{2-} ions).

Naming Compounds

- Early compound names were based on physical properties, origin, or use (e.g., milk of magnesia, laughing gas, baking soda).
- With over **66 million compounds** today, a systematic naming system is essential.

❖ Systematic Naming:

1. Chemists developed **global rules** for naming substances to facilitate communication and labeling.

➤ Organic vs. Inorganic Compounds:

1. Organic compounds: Contain carbon, usually combined with hydrogen, oxygen, nitrogen, or sulfur.

2. Inorganic compounds: All other compounds, including:

- Carbon monoxide (**CO**), carbon dioxide (**CO₂**), carbon disulfide (**CS₂**).
- Compounds with cyanide (**CN⁻**), carbonate (**CO₃²⁻**), and bicarbonate (**HCO₃⁻**).

➤ Categories of Inorganic Compounds:

- 1. Ionic compounds:** Composed of cations and anions.
- 2. Molecular compounds:** Formed by nonmetals.
- 3. Acids and bases:** Compounds that donate or accept protons or hydroxide ions.
- 4. Hydrates:** Compounds containing water molecules in their structure.

This framework simplifies the naming process for both organic and inorganic compounds.

Naming Compounds

❑ Naming Ionic Compounds

• Ionic Compound Basics:

- Composed of **cations** (positive ions) and **anions** (negative ions).
- Most cations are derived from **metal atoms**, with the exception of the **ammonium ion (NH_4^+)**.

• Naming Cations:

- Cations are named after the **metal elements** they are derived from.

• Naming Binary Ionic Compounds

- **Binary ionic compounds** consist of **two elements**: a metal cation and a nonmetal anion.

• Naming Rules:

- Name **the metal cation** first.
- Name the **nonmetal anion** by taking the root of the element's name and adding **"-ide"**.

• Examples:

- **NaCl**: Sodium chloride (sodium = cation, chloride = anion derived from chlorine).
- **MgO**: Magnesium oxide.
- **CaF₂**: Calcium fluoride.

Element		Name of Cation	
Na	sodium	Na^+	sodium ion (or sodium cation)
K	potassium	K^+	potassium ion (or potassium cation)
Mg	magnesium	Mg^{2+}	magnesium ion (or magnesium cation)
Al	aluminum	Al^{3+}	aluminum ion (or aluminum cation)

Table 2.2

The “-ide” Nomenclature of Some Common Monatomic Anions According to Their Positions In the Periodic Table

Group 4A	Group 5A	Group 6A	Group 7A
C carbide (C^{4-})*	N nitride (N^{3-})	O oxide (O^{2-})	F fluoride (F^-)
Si silicide (Si^{4-})	P phosphide (P^{3-})	S sulfide (S^{2-})	Cl chloride (Cl^-)
		Se selenide (Se^{2-})	Br bromide (Br^-)
		Te telluride (Te^{2-})	I iodide (I^-)

*The word “carbide” is also used for the anion C_2^{2-} .

Naming Compounds

- **Naming Ionic Compounds:**
- **Binary Compounds:**
 - ✓ Made of two elements: a metal cation and a nonmetal anion.
 - ✓ **Examples:**
 - **KBr:** Potassium bromide.
 - **ZnI₂:** Zinc iodide.
 - **Al₂O₃:** Aluminum oxide.
- **"-ide" Nomenclature:**
 - Used for naming monatomic anions.
 - Examples:
 - Chloride (**Cl⁻**), oxide (**O²⁻**), bromide (**Br⁻**).
 - Also applies to specific polyatomic anions:
 - **Hydroxide (OH⁻).**
 - **Cyanide (CN⁻).**
- **Ternary Compounds:**
 - Composed of **three elements**.
 - Examples:
 - **LiOH:** Lithium hydroxide.
 - **KCN:** Potassium cyanide.

Element		Name of Cation	
Na	sodium	Na ⁺	sodium ion (or sodium cation)
K	potassium	K ⁺	potassium ion (or potassium cation)
Mg	magnesium	Mg ²⁺	magnesium ion (or magnesium cation)
Al	aluminum	Al ³⁺	aluminum ion (or aluminum cation)

Table 2.2

The “-Ide” Nomenclature of Some Common Monatomic Anions According to Their Positions In the Periodic Table

Group 4A	Group 5A	Group 6A	Group 7A
C carbide (C ⁴⁻)*	N nitride (N ³⁻)	O oxide (O ²⁻)	F fluoride (F ⁻)
Si silicide (Si ⁴⁻)	P phosphide (P ³⁻)	S sulfide (S ²⁻)	Cl chloride (Cl ⁻)
		Se selenide (Se ²⁻)	Br bromide (Br ⁻)
		Te telluride (Te ²⁻)	I iodide (I ⁻)

*The word “carbide” is also used for the anion C₂²⁻.

Naming Compounds

- **Transition Metals with Multiple Cations:**
 - Some metals can form more than one cation (e.g., Fe^{2+} and Fe^{3+} for iron).
 - Older naming system:
 - **"-ous"** for the cation with fewer positive charges: Ferrous (Fe^{2+}).
 - **"-ic"** for the cation with more positive charges: Ferric (Fe^{3+}).
- **Limitations of "-ous" and "-ic" Naming and the Stock System:**
- **Limitations of "-ous" and "-ic" Naming:**
 - Does not indicate the **specific charge** of the cation.
 - Example: Ferric ion (Fe^{3+}) and cupric ion (Cu^{2+}).
 - Can only name **two different cations** for an element, while some elements (e.g., manganese) can have three or more charges.
- **Stock System:**
 - Uses **Roman numerals** to indicate the **exact charge** of the cation.
 - Roman numeral system:
 - **I**: One positive charge.
 - **II**: Two positive charges.
 - **III**: Three positive charges, and so on.
- **Examples with Manganese (Mn):**
 - Mn^{2+} : Manganese(II).
 - Mn^{3+} : Manganese(III).
 - Mn^{4+} : Manganese(IV).
 - Mn^{7+} : Manganese(VII).
- The **Stock system** has become the standard for clearly and accurately naming cations with variable charges.

Problems

Name the following compounds: (a) $\text{Fe}(\text{NO}_3)_2$, (b) Na_2HPO_4 , and (c) $(\text{NH}_4)_2\text{SO}_3$.

Write chemical formulas for the following compounds: (a) mercury(I) nitrate, (b) cesium oxide, and (c) strontium nitride.

Table 2.3

Names and Formulas of Some Common Inorganic Cations and Anions

Cation	Anion
aluminum (Al^{3+})	bromide (Br^-)
ammonium (NH_4^+)	carbonate (CO_3^{2-})
barium (Ba^{2+})	chlorate (ClO_3^-)
cadmium (Cd^{2+})	chloride (Cl^-)
calcium (Ca^{2+})	chromate (CrO_4^{2-})
cesium (Cs^+)	cyanide (CN^-)
chromium(III) or chromic (Cr^{3+})	dichromate ($\text{Cr}_2\text{O}_7^{2-}$)
cobalt(II) or cobaltous (Co^{2+})	dihydrogen phosphate (H_2PO_4^-)
copper(I) or cuprous (Cu^+)	fluoride (F^-)
copper(II) or cupric (Cu^{2+})	hydride (H^-)
hydrogen (H^+)	hydrogen carbonate or bicarbonate (HCO_3^-)
iron(II) or ferrous (Fe^{2+})	hydrogen phosphate (HPO_4^{2-})
iron(III) or ferric (Fe^{3+})	hydrogen sulfate or bisulfate (HSO_4^-)
lead(II) or plumbous (Pb^{2+})	hydroxide (OH^-)
lithium (Li^+)	iodide (I^-)
magnesium (Mg^{2+})	nitrate (NO_3^-)
manganese(II) or manganous (Mn^{2+})	nitride (N^{3-})
mercury(I) or mercurous (Hg_2^{2+})*	nitrite (NO_2^-)
mercury(II) or mercuric (Hg^{2+})	oxide (O^{2-})
potassium (K^+)	permanganate (MnO_4^-)
rubidium (Rb^+)	peroxide (O_2^{2-})
silver (Ag^+)	phosphate (PO_4^{3-})
sodium (Na^+)	sulfate (SO_4^{2-})
strontium (Sr^{2+})	sulfide (S^{2-})
tin(II) or stannous (Sn^{2+})	sulfite (SO_3^{2-})
zinc (Zn^{2+})	thiocyanate (SCN^-)

*Mercury(I) exists as a pair as shown.

Naming Compounds

- **Naming Binary Molecular Compounds:**
- **Composition:**
 - Typically composed of **nonmetallic elements**.
- **Naming Rules:**
 - Name the **first element** as it appears in the formula.
 - Name the **second element** by adding **"-ide"** to the root of its name.
- **Examples:**
 - ✓ **HCl**: Hydrogen chloride.
 - ✓ **HBr**: Hydrogen bromide.
 - ✓ **SiC**: Silicon carbide
- **Naming Molecular Compounds Using Greek Prefixes:**
- **Greek Prefixes:**
 - Used to indicate the number of atoms of each element in a compound.
- **Examples:**
 - ✓ **CO**: Carbon monoxide.
 - ✓ **CO₂**: Carbon dioxide.
 - ✓ **SO₂**: Sulfur dioxide.
 - ✓ **SO₃**: Sulfur trioxide.
 - ✓ **NO₂**: Nitrogen dioxide.
 - ✓ **N₂O₄**: Dinitrogen tetroxide.

Greek Prefixes Used in Naming Molecular Compounds

Prefix	Meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Naming Compounds

- **Naming Guidelines:**
 - ✓ "Mono-" is usually omitted for the first element.
 - Example: PCl_3 is phosphorus trichloride, not monophosphorus trichloride.
 - ✓ For **oxides**, the ending "a" in the prefix is sometimes dropped:
 - Example: N_2O_4 is dinitrogen tetroxide, not dinitrogen tetraoxide.
- **Exceptions (Hydrogen Compounds):**
 - Compounds containing hydrogen often have **nonsystematic names**:
 - ✓ B_2H_6 : Diborane.
 - ✓ CH_4 : Methane.
 - ✓ SiH_4 : Silane.
 - ✓ NH_3 : Ammonia.
 - ✓ PH_3 : Phosphine.
 - ✓ H_2O : Water.
 - ✓ H_2S : Hydrogen sulfide.
 - **Order of elements** can vary:
 - In water (H_2O) and hydrogen sulfide (H_2S), **H** is written first.
 - In other compounds (e.g., CH_4), **H** appears last.
- **Writing Formulas:**
 - The formula follows the order of elements in the name.
 - Example: **Arsenic trifluoride (AsF_3)** contains 1 As atom and 3 F atoms.

Problems

Name the following molecular compounds:

(a) PBr_5

(b) As_2O_5

(c) NF_3

(d) Cl_2O_7

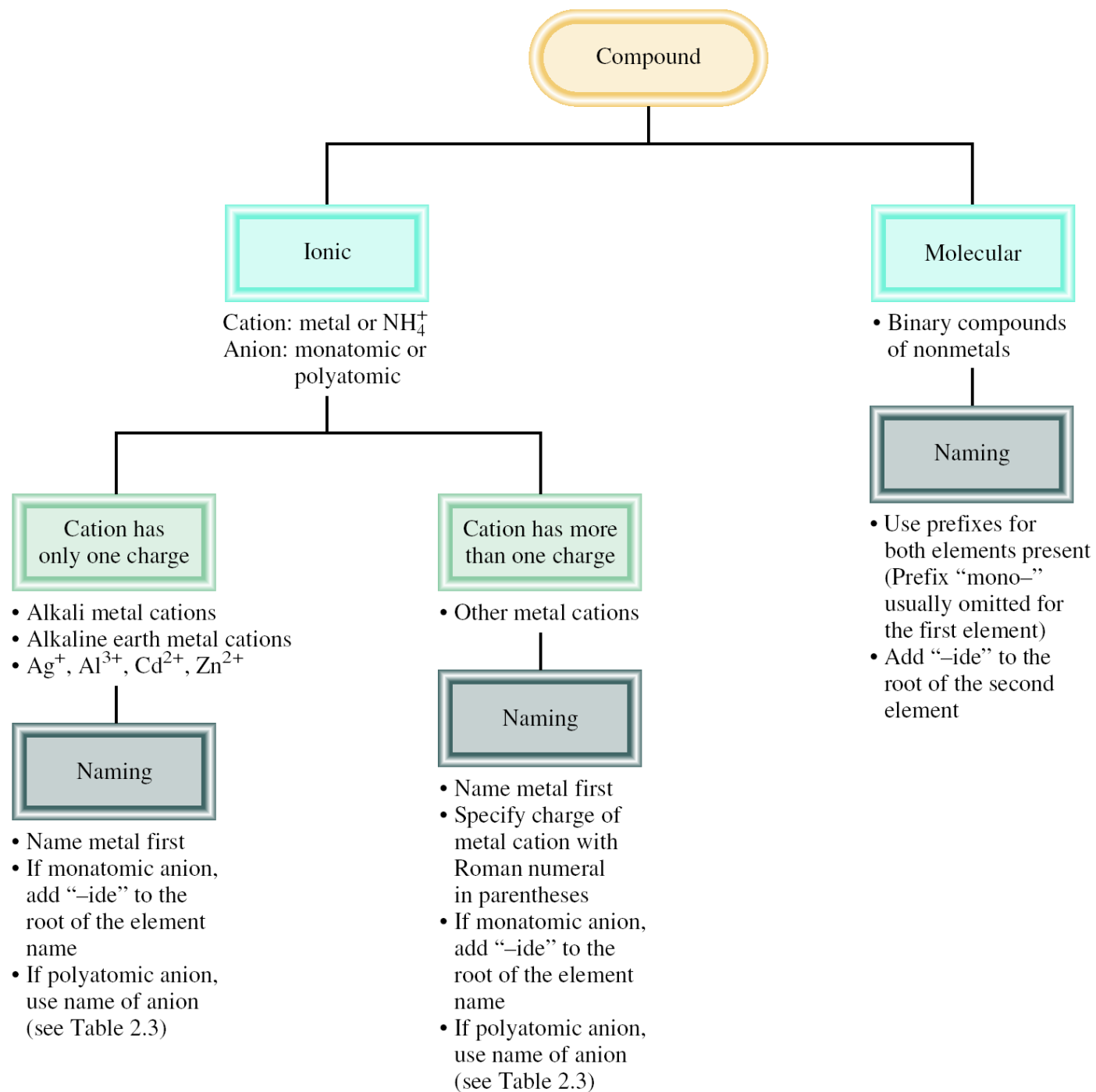
Write chemical formulas for the following molecular compounds:

(a) bromine trifluoride

(b) diboron trioxide.

(c) sulfur tetrafluoride

(d) dinitrogen pentoxide.



Acids and Bases

- **Naming Acids**

- ✓ Acid is a substance that yields **hydrogen ions (H^+)** when dissolved in water.
- ✓ **Formula:** Contains hydrogen atoms and an anionic group.

- **Naming Binary Acids (Anion ends in "-ide"):**

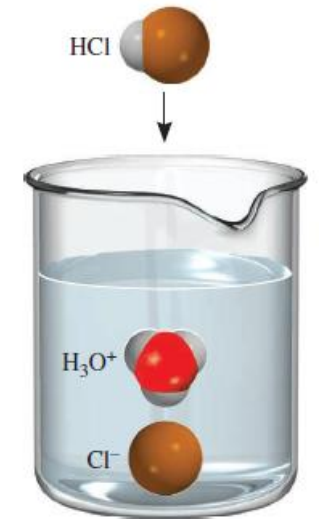
- Prefix "**hydro-**" and suffix "**-ic**" added to the root of the anion.
- Example:
 - ✓ **HCl (gas):** Hydrogen chloride.
 - ✓ **HCl (aqueous):** Hydrochloric acid.

- **Naming Oxoacids:**

- ✓ Contain hydrogen, oxygen, and a central element.
- ✓ Formula: **H** written first, followed by the central element and then oxygen.

- **Reference acids:**

- **H₂CO₃:** Carbonic acid.
- **HClO₃:** Chloric acid.
- **HNO₃:** Nitric acid.
- **H₃PO₄:** Phosphoric acid.
- **H₂SO₄:** Sulfuric acid.



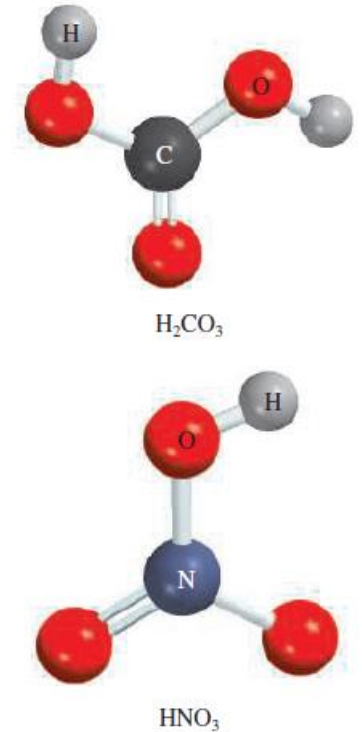
When dissolved in water, the HCl molecule is converted to the H^+ and Cl^- ions. The H^+ ion is associated with one or more water molecules, and is usually represented as H_3O^+ .

Table 2.5 Some Simple Acids

Acid	Corresponding Anion
HF (hydrofluoric acid)	F^- (fluoride)
HCl (hydrochloric acid)	Cl^- (chloride)
HBr (hydrobromic acid)	Br^- (bromide)
HI (hydroiodic acid)	I^- (iodide)
HCN (hydrocyanic acid)	CN^- (cyanide)
H ₂ S (hydrosulfuric acid)	S^{2-} (sulfide)

Acids and Bases

- **Naming Acids**
- **Modifications to Oxoacid Names:**
- **Addition** of **one O** atom to the “**-ic**” acid: The acid is called “**per . . . -ic**” acid. Thus, adding an O atom to **HClO₃** changes chloric acid to perchloric acid, **HClO₄**.
- **Removal** of one **O** atom from the “**-ic**” acid: The acid is called “**-ous**” acid. Thus, nitric acid, **HNO₃**, becomes **nitrous acid**, **HNO₂**.
- **Removal** of **two O** atoms from the “**-ic**” acid: The acid is called “**hypo . . . -ous**” acid. Thus, when **HBrO₃** is converted to **HBrO**, the acid is called **hypobromous acid**.



Acids and Bases

- **Rules for Naming Oxoanions (Anions of Oxoacids)**

- **“-ate” Anions:**

- When all hydrogen ions (H^+) are removed from an “-ic” acid, the anion’s name ends with “-ate.”
- Example: The anion CO_3^{2-} derived from H_2CO_3 is called **carbonate**.

- **“-ite” Anions:**

- When all hydrogen ions (H^+) are removed from an “-ous” acid, the anion’s name ends with “-ite.”
- Example: The anion ClO_2^- derived from HClO_2 is called **chlorite**.

- **Anions with Partial Hydrogen Removal:**

- If one or more (but not all) hydrogen ions (H^+) are removed, the name of the anion must indicate the number of hydrogen ions still present.
- Example: Anions derived from **phosphoric acid (H_3PO_4)**:
 - H_2PO_4^- : **Dihydrogen phosphate**
 - HPO_4^{2-} : **Hydrogen phosphate**
 - PO_4^{3-} : **Phosphate**

- **Note:**

- The prefix "**mono-**" is typically omitted when there is only one hydrogen atom (H^+) in the anion.

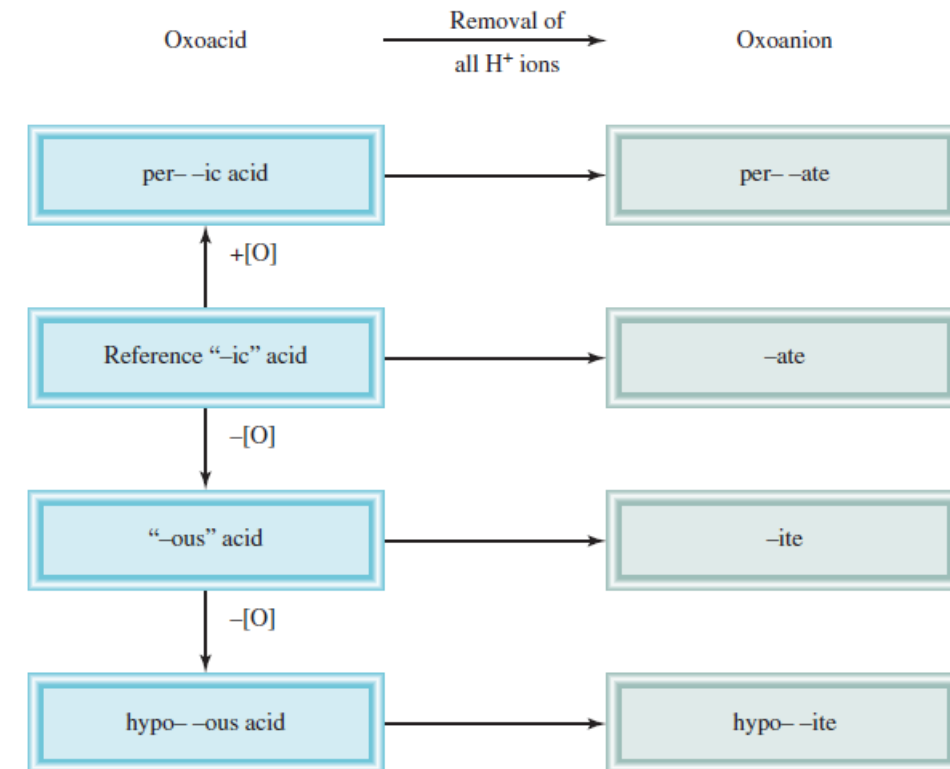


Table 2.6 Names of Oxoacids and Oxoanions That Contain Chlorine

Acid	Corresponding Anion
HClO_4 (perchloric acid)	ClO_4^- (perchlorate)
HClO_3 (chloric acid)	ClO_3^- (chlorate)
HClO_2 (chlorous acid)	ClO_2^- (chlorite)
HClO (hypochlorous acid)	ClO^- (hypochlorite)

Name the following oxoacid and oxoanions: (a) H_2SO_3 , a very unstable acid formed when $\text{SO}_2(g)$ reacts with water, (b) H_2AsO_4 , once used to control ticks and lice on livestock, and (c) SeO_3^{2-} , used to manufacture colorless glass. H_3AsO_4 is arsenic acid, and H_2SeO_4 is selenic acid.

Acids and Bases

- **Naming Bases**
- A base is a substance that yields hydroxide ions (OH^-) when dissolved in water.
- **Examples of Bases:**
- **NaOH**: Sodium hydroxide
- **KOH**: Potassium hydroxide
- **Ba(OH)₂**: Barium hydroxide
- **Ammonia (NH₃)**:
- **Ammonia** is a molecular compound that exists as a gas or a pure liquid.
- Despite not containing hydroxide ions in its structure, **it is classified as a base** because it reacts with water to produce hydroxide ions:
$$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$$
- This ability to yield hydroxide ions in solution makes ammonia a base.

Hydrates

➤ **Hydrates** are compounds that have a specific number of water molecules associated with each formula unit.

- **Naming Hydrates:**

- The number of water molecules is indicated using Greek prefixes in the compound's name, followed by the word "hydrate."

➤ **Examples:**

- $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$: Copper(II) sulfate pentahydrate
- $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$: Barium chloride dihydrate
- $\text{LiCl} \cdot \text{H}_2\text{O}$: Lithium chloride monohydrate
- $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$: Magnesium sulfate heptahydrate
- $\text{Sr}(\text{NO}_3)_2 \cdot 4\text{H}_2\text{O}$: Strontium nitrate tetrahydrate



$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

CuSO_4

- **Anhydrous Compounds:**

- When the water molecules are removed (typically by heating), the compound becomes **anhydrous**.

➤ **Example:** Heating $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ results in CuSO_4 , called anhydrous copper(II) sulfate.

- **Note on "Anhydrous":**

- "Anhydrous" means that the compound no longer contains water molecules in its structure.

- This systematic naming ensures clarity when identifying and discussing hydrated compounds.

Introduction to Organic Compounds

- **Organic Compounds: Hydrocarbons and Functional Groups**

- **Hydrocarbons:**

- The simplest type of organic compounds, consisting only of carbon and hydrogen atoms.
- **Uses:**
 - Fuels for domestic and industrial heating.
 - Generating electricity.
 - Powering internal combustion engines.
 - Starting materials in the chemical industry.

- **Alkanes:**

- A class of hydrocarbons with only single bonds between carbon atoms.
- **Straight-Chain Alkanes:**
 - Names of straight-chain alkanes end with **-ane**.
 - **Examples:**
 - **Methane** (CH_4)
 - **Ethane** (C_2H_6)
 - **Propane** (C_3H_8)
 - Starting with C_5H_{12} , Greek prefixes indicate the number of carbon atoms.

Greek Prefixes Used in Naming Molecular Compounds

Prefix	Meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Introduction to Organic Compounds

- **Functional Groups:**
 - Functional groups largely determine the chemistry of organic compounds.
 - **Examples of Functional Groups:**
 - **Hydroxyl Group (-OH):** Replacing an H atom in methane forms alcohols.
 - **Amino Group (-NH₂):** Replacing an H atom creates amines.
 - **Carboxyl Group (-COOH):** Replacing an H atom forms carboxylic acids.
- **Resulting Molecules:**
 - Substituting functional groups into methane (CH₄) produces:
 - **Methanol (CH₃OH):** Hydroxyl group (-OH).
 - **Methylamine (CH₃NH₂):** Amino group (-NH₂).
 - **Methanoic Acid (HCOOH):** Carboxyl group (-COOH).
- This classification and naming system aids in identifying hydrocarbons and understanding their transformations into functionalized organic compounds.

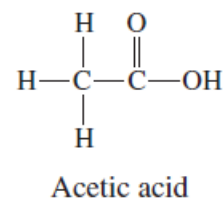
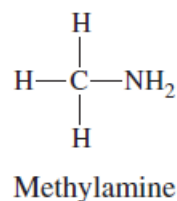
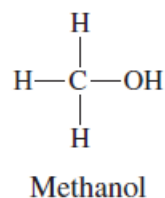


Table 2.8 The First Ten Straight-Chain Alkanes		
Name	Formula	Molecular Model
Methane	CH ₄	
Ethane	C ₂ H ₆	
Propane	C ₃ H ₈	
Butane	C ₄ H ₁₀	
Pentane	C ₅ H ₁₂	
Hexane	C ₆ H ₁₄	
Heptane	C ₇ H ₁₆	
Octane	C ₈ H ₁₈	
Nonane	C ₉ H ₂₀	
Decane	C ₁₀ H ₂₂	

THANK YOU.



ANY
QUESTIONS?

dreamstime

