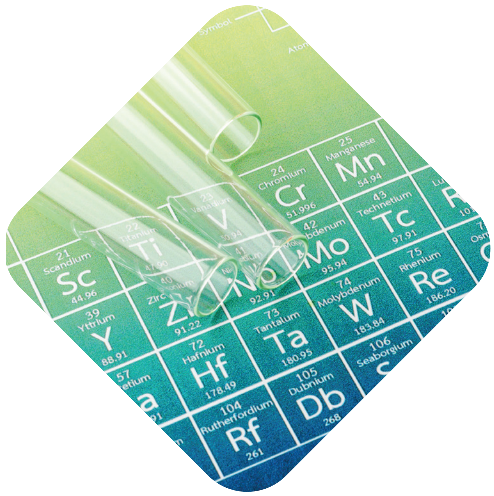
MINISTRY OF EDUCATION AND TECHNICAL EDUCATION
وزارة التربية والتعليم والتعليم الفني

Chemistry

For 2nd secondary - First Term

2025-2026



Chemistry

For 2nd secondary - First Term 2025-2026

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Introduction

In the context of developing education to keep pace with global and local changes, and as a continuation of the diligent efforts made by the Ministry of Education to elevate the level of curriculum content and link it to society and the environment, the Minister of Education has commissioned a reevaluation and review of the scientific content of the chemistry subject for the second secondary grade.

The committee assigned to make the necessary amendments and additions has led to:

(1) Eliminating repetition and unjustified padding, excluding parts that students have previously studied, and rephrasing some parts of the book in a logical, sequential, and organized manner.

(2) Adding some concepts and applications to keep pace with modern scientific trends.

(3) Linking study topics to daily life, their environmental impacts, and industrial applications.

(4) Focusing on mathematical treatments to help students understand certain topics and strengthen their intellectual skills..

(5) Preparing illustrative figures and employing them to support scientific concepts.

(6) Defining the expected learning objectives for each chapter in its introduction to provide students and teachers with clear indicators of achievement..

(7) Ensuring diversity in assessment to measure different levels of learning.

The book, in its current form, contains three integrated and interconnected chapters that align with the content of chemistry books in global curricula and include useful industrial and environmental applications.
It demonstrates a clear focus on developing students' understanding, analytical, and innovative skills, in alignment with the national standards set by the ministry for the development of the chemistry curriculum.

We hope that this book, in its new form, will serve as a valuable source of knowledge in the field of chemistry, achieve its intended objectives, and provide effective support to our students, to whom we wish success and prosperity.

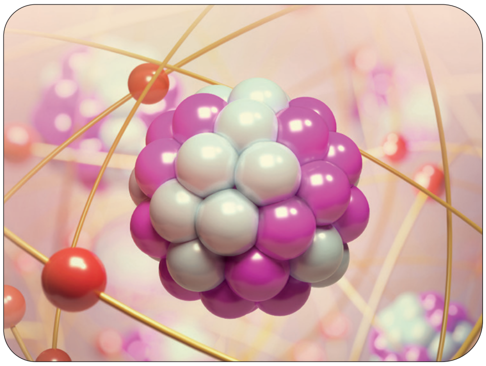
Development Committee

Contents of the Book

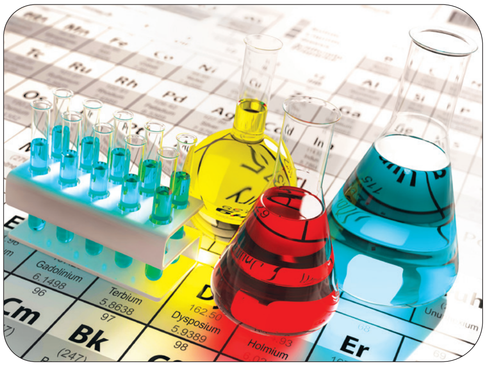
First Term

Chapter One

Chapter Two



Chapter Three



Chapter One

Chemical Calculation

Objectives

- By the end of this chapter the student should be able to:

1- Define the mole.

2- Calculate the molar masses of some compounds.

3- Apply Avogadro's hypothesis in chemical calculations for gas reactions.

4- Determine the limiting reactant of a reaction.

5- Differentiate between empirical and molecular formulas.

6- Deduce the molecular formula for some compounds.

7- Calculate the volumes of reactant or product gases in a chemical reaction.

8- Calculate the molar concentration (molarity) of some solutions.

The mole

The amount of a chemical substance is expressed in the International System of Units (SI) in in mole unit, and one mole of any substance contains a fixed number of particles (molecules, atoms or ions) called **Avogadro's number** which equals 6.02×10^{23}

Number of particles = Number of moles of particles × Avogadro's number

Example

Calculate the number of atoms of hydrogen in 2 mol of methane CH_4

Answer

$\because$ 1 mol of CH_4 contains 4 mol atom of H

$\therefore$ 2 mol of CH_4 contains 8 mol atom of H

The number of hydrogen atoms $= 6.02 \times 10^{23} \times 8$

$= 48.16 \times 10^{23}$ atom

And **the molar mass** represents the gram-relative atomic mass of the substances or the gram molecular mass, measured in g/mol, and equals the sum of the gram-relative atomic masses of the substances constituting the molecule.

Application

1 Calculate the molar mass of water H_2O in terms of the gram-relative atomic masses of the elements which constitute one molecule of it. **[H = 1 , O = 16]**

Answer

The molar mass of $H_2O = 16 + (2 \times 1) = 18$ g/mol

2 Calculate the molar mass of hydrated blue copper sulphate crystals $CuSO_4 \cdot 5H_2O$

[Cu = 63.5 , S = 32 , O = 16 , H = 1]

Answer

$$\underset{63.5}{Cu} \quad \underset{32}{S} \quad \underset{(16 \times 4)}{O_4} \quad \cdot \quad \underset{5[(1 \times 2) + 16]}{5H_2O}$$

$$63.5 + 32 + (16 \times 4) + 5\,[(1 \times 2) + 16] = 249.5 \text{ g/mol}$$

The relation between the mass of substance, its number of moles, and its molar mass is given by the following relation:

$$\text{Number of moles (mol)} = \frac{\text{Mass of the substance (g)}}{\text{Molar mass (g/mol)}}$$

Example

A sample of sodium carbonate has mass of 265 g

Given the relative atomic masses of the constituent elements [**Na = 23 , C = 12 , O = 16**],

Calculate:

1. Number of moles in the sample.
2. Number of moles of sodium ions in the sample.
3. Number of sodium ions produced from dissolving the sample in water.

Answer

1. Molar mass of $Na_2CO_3 = (2 \times 23) + 12 + (3 \times 16) = 106$ g/mol

$$\text{Number of moles} = \frac{\text{Mass of the substance}}{\text{Molar mass}} = \frac{265}{106} = 2.5 \text{ mol}$$

2. $Na_2CO_{3(s)} \xrightarrow{\text{water}} 2Na^+_{(aq)} + CO^{2-}_{3(aq)}$

   1 mol   2 mol

   ∵ Since each 1 mol of Na_2CO_3 contains 2 mol of Na^+ ions

   ∴ Therefore, the number of moles of Na^+ ions in the sample = $2 \times 2.5 = 5$ mol ion

3. ∵ Since the number of Na^+ ions = number of moles of Na^+ × Avogadro's number

   ∴ Therefore, the number of Na^+ ions = $5 \times 6.02 \times 10^{23} = 30.1 \times 10^{23}$ ions

Molar volume of gases

Avogadro's law states that one mole of any gas occupies the same volume under the same conditions of pressure and temperature. Later measurements revealed that one mole of any gas occupies a volume of 22.4 L under standard conditions (STP), namely a temperature of 0°C (273 K) and a normal atmospheric pressure of 1 atm (760 mmHg) or occupies a volume of 24 L at (rtp) room temperature pressure which is a temperature of 25°C (298 K) and normal atmospheric pressure.

Avogadro's postulate states that equal volumes of different gases under the same conditions of pressure and temperature contain the same number of molecules.

The same volume of any gas at STP contains the same number of molecules.

The following table shows the relation: between the number of moles, volume, and number of molecules for some gases **(at STP)** and **(at rtp):**

Gas	O_2	N_2	H_2
Number of moles	1 mol	1 mol	1 mol
Volume (at STP)	22.4 L	22.4 L	22.4 L
Volume (at rtp)	24 L	24 L	24 L
Number of molecules	6.02×10^{23} molecule	6.02×10^{23} molecule	6.02×10^{23} molecule
Molar mass	32 g/mol	28 g/mol	2 g/mol

Example

Calculate the volume occupied by 1.6 g of methane gas CH_4 at:

1 STP

2 rtp

[C = 12 , H = 1]

Answer

molar mass of $CH_4 = 12 + (4 \times 1) = 16$ g/mol

Number of moles of $CH_4 = \frac{1.6}{16} = 0.1$ mol

1 Volume of gas (at STP) = $0.1 \times 22.4 = 2.24$ L

2 Volume of gas (at rtp) = $0.1 \times 24 = 2.4$ L

Chemical calculation in terms of the balanced symbolic equation

The balanced symbolic equation is used to calculate or determine each of:

1. Masses of substances.
2. Number of moles.
3. Number of particles.
4. Volumes of gases.
5. The limiting reactant of the reaction.

Application (1)

The acidified water is electrolyzed according to the balanced equation:

$$2H_2O_{(\ell)} \longrightarrow 2H_{2(g)} + O_{2(g)}$$

Calculate: **[H = 1 , O = 16]**

1. Number of water moles needed to produce 2.5 mole of oxygen.
2. Mass of water needed to produce 2 g of hydrogen.
3. Number of oxygen molecules produced from the electrolysis of 1 mol of water.

Answer

1.

$$\begin{array}{ccc} 2H_2O_{(\ell)} & \longrightarrow & O_{2(g)} \\ 2\text{ mol} & & 1\text{ mol} \\ ?\text{ mol} & & 2.5\text{ mol} \end{array}$$

Number of water moles = 2.5 × 2 = 5 mol

2.

$$\begin{array}{ccc} 2H_2O_{(\ell)} & \longrightarrow & 2H_{2(g)} \\ 2\,(2 + 16) = 36\text{ g} & & 2\,(1 \times 2) = 4\text{g} \\ ?\text{ g} & & 2\text{ g} \end{array}$$

Mass of water $= \frac{36 \times 2}{4} = 18$ g

3.

$$\begin{array}{ccc} 2H_2O_{(\ell)} & \longrightarrow & O_{2(g)} \\ 2\text{ mol} & & 1\text{ mol} \\ 1\text{ mol} & & ?\text{ mol} \end{array}$$

Number of O_2 moles $= \frac{1 \times 1}{2} = 0.5$ mol

Number of O_2 molecules $= 0.5 \times 6.02 \times 10^{23}$

$= 3.01 \times 10^{23}$ molecule

Application (2)

Potassium nitrate decomopses upon heating, according to the following symbolic equation:

$$2KNO_{3(s)} \xrightarrow{\Delta} 2KNO_{2(s)} + O_{2(g)}$$

Calculate the volume of oxygen gas produced from the decomposition of 5 mol of potassium nitrate at:

1 STP 2 rtp

$$2KNO_{3(s)} \longrightarrow O_{2(g)}$$

$2KNO_{3(s)}$	$O_{2(g)}$
2 mol	1 mol
5 mol	? mol

Number of O_2 moles produced $= \frac{5}{2} = 2.5$ mol

1 Volume of O_2 gas (at STP) $= 2.5 \times 22.4 = 56$ L

2 Volume of O_2 gas (at rtp)$= 2.5 \times 24 = 60$ L

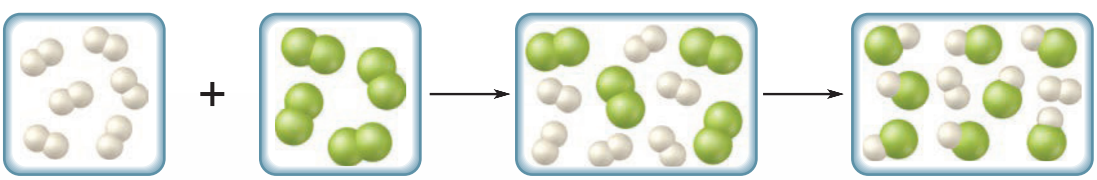
The limiting reactant of the reaction

Hydrogen gas reacts with fluorine gas, according to the following balanced equation:

$$H_{2(g)} + F_{2(g)} \longrightarrow 2HF_{(g)}$$

1 mol 1 mol 2 mol

What happens when 6 mol of hydrogen gas is added to 4 mol of fluorine gas in the reaction vessel, As shown in the following figure:



6 mol of H_2 | 4 mol of F_2 | Mixture of reactants | The result product of the reaction is 8 mol of HF
The remaining substance is 2 mol of H_2

The entire amount of fluorine (4 mol) is completely consumed in the reaction with 4 mol of hydrogen to form 8 mol of hydrogen fluoride, 2 mol of hydrogen gas remain unreacted (excess amount).

Fluorine is described in this reaction as **the limiting reactant**, because its amount is entirely consumed in the reaction, and it determines the amount of the formed product.

Calculation of the limiting reactant

The limiting reactant is determined by calculating the ratio of the number of moles present in the reaction mixture to the number of moles in the balanced chemical equation for each reactant . The reactant with the smaller ratio is the limiting reactant. It is the substance that is completely consumed first, causing the reaction to stop even if other reactants are present in excess.

Example

from the reaction: $2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(\ell)}$

Clarify with chemical calculations:

1. The limiting reactant when mixing 14.4 L of oxygen gas with 24 L of hydrogen gas (at rtp) under suitable conditions for the reaction.
2. The excess mass of the other reactant.

[H = 1 , O = 16]

Answer

1. Number of moles $= \dfrac{\text{Volume of gas}}{24}$

Number of given moles of $O_2 = \dfrac{14.4}{24} = 0.6$ mol

Number of given moles of $H_2 = \dfrac{24}{24} = 1$ mol

Ratio of the number of moles of each reactant =

$$\frac{\text{Number of moles given}}{\text{Number of moles in the balanced symbolic equation}}$$

Ratio of the number of moles of $O_2 = \dfrac{0.6}{1} = 0.6$

Ratio of the number of moles of $H_2 = \dfrac{1}{2} = 0.5$

$\because$ The ratio of the number of moles of H_2 is less.

$\therefore$ The limiting reactant for the reaction is hydrogen.

2 $2 \text{ mol } H_2 \xrightarrow[\text{completely with}]{\text{It reacts}} 1 \text{ mol } O_2$ «From the balanced symbolic equation»

$2 \times 1 \times 2 = 4 \text{ g} \longrightarrow 1 \times 16 \times 2 = 32 \text{ g}$

$1 \times 1 \times 2 = 2 \text{ g}$ $\quad$? g

Mass of O_2 reactant $= \frac{32 \times 2}{4} = 16$ g

Mass of O_2 given $= 0.6 \times 32 = 19.2$ g

Excess mass of O_2 (remaining unreacted) $= 19.2 - 16 = 3.2$ g

Calculating chemical formulas

Chemical formulas of compounds are classified in terms of the type of information they provide, into three main types, which are:

- The empirical formula.
- The molecular formula.
- The structural formula.

The empirical formula expresses the simplest integer numerical ratio of the atoms or ions of the elements composing the molecule.

The molecular formula expresses the types and numbers of the atoms or ions that compose the molecule.

the structural formula expresses the types and numbers of the atoms of the elements in the molecule and the way of their bonding together.

The following table shows the empirical formula and molecular formula of some compounds:

Compound	Empirical formula	Molecular formula
Ethane	CH_3	C_2H_6
Hydrogen peroxide	HO	H_2O_2
Hydrazine	NH_2	N_2H_4

It is concluded **that the empirical formula** of an unknown compound can be determined based on the percentage of the elements constituting the compound, considering this percentage represents the masses of each element present in every 100 g of the compound. and with the help of experimental results, **this is done by calculating:**

- The number of moles of atoms of each element in the compound.
- The ratio of the number of moles of atoms of each element in the compound molecule divided by the smallest value of the number of moles, from which the empirical formula of the compound can be determined.

Application

Calculate the empirical formula for compound (X)
which consists of 85.63% C, and 14.37% H **[C = 12 , H = 1]**

Answer Outline

Steps	H	C
Number of moles of atoms of each element	$\frac{14.37}{1} = 14.37$ mol	$\frac{85.63}{12} = 7.136$ mol
Ratio of the number of moles of atoms of each element	$\frac{14.37}{7.136} = 2$	$\frac{7.136}{7.136} = 1$

∴ The empirical formula of the compound (X) : CH_2

The **molecular formula** of compounds is calculated based on the number of units of the empirical formula (n) which is calculated as follows :

$$\text{Number of units of the empirical formula (n)} = \frac{\text{The molar mass of the compound}}{\text{The molar mass of the empirical formula}}$$

If the molar mass of the compound (X) equals 28 g/mol

And the molar mass of the empirical formula (CH_2) equals 14 g/mol

Then: $n = \frac{28}{14} = 2$

Its molecular formula is: $(CH_2)_2 = C_2H_4$

Example (1)

A hydrocarbon compound in which the mass percentage of carbon is 83.7 %

and its molar mass 86 g/mol **[C = 12 , H = 1]**

Calculate the molecular formula of this compound

Answer

The mass percentage of hydrogen in the compound = 100 % – 83.7 % = 16.3%

The number of moles of C atoms in the compound = $\frac{83.7}{12}$ = 6.975 mol

The number of moles of H atoms in the compound = $\frac{16.3}{1}$ = 16.3 mol

The ratio of the number of moles of C atoms in the compound = $\frac{6.975}{6.975}$ = 1

The ratio of the number of moles of H in the compound = $\frac{16.3}{6.975}$ = 2.337

The simplest integer numerical ratio of H , C atoms in the empirical formula by multiplying (× 3): C_3H_7

The molar mass from the empirical formula = (3 × 12) + 7 = 43 g/mol

The number of units of the empirical formula (n) = $\frac{86}{43}$ = 2

The molecular formula of the compound : $(C_3H_7)_2 = C_6H_{14}$

Example (2)

A hydrocarbon with molecular formula C_xH_y and molar mass 58 g/mol produces, upon complete combustion in an abundance of oxygen gas, an amount of carbon dioxide gas with a mass of 88 g [CO_2 = 44 g/mol] and an amount of water vapor with a mass of 45 g [H_2O = 18 g/mol] Calculate the value of both x and y **[C = 12 , H = 1]**

Answer

∵ Carbon atoms in the hydrocarbon compound convert into molecules of carbon dioxide CO_2

∴ Mass of CO_2 is used to calculate the mass of C in the hydrocarbon.

Mass of carbon in CO_2 = the mass of CO_2 × $\frac{\text{The molar mass of C}}{\text{The molar mass of } CO_2}$

Mass of carbon in CO_2 = $\frac{12}{44}$ × 88 = 24 g

∵ The hydrogen atoms of the hydrocarbon compound convert into molecules of water vapor H_2O

∴ Mass of H_2O is used to calculate the mass of H in the hydrocarbon.

Mass of hydrogen in H_2O = Mass of H_2O × $\frac{\text{Molar mass of } H_2}{\text{Molar mass of } H_2O}$

$$= 45 \times \frac{2}{18} = 5 \text{ g}$$

The number of moles of C in the hydrocarbon $= \frac{24}{12} = 2$ mol

The number of moles of H in the hydrocarbon $= \frac{5}{1} = 5$ mol

The ratio of the number of moles of C atoms in the hydrocarbon $= \frac{2}{2} = 1$

The ratio of the number of moles of H atoms in the hydrocarbon $= \frac{5}{2} = 2.5$

The simplest integer numerical ratio of atoms C , H in the empirical formula of the hydrocarbon (by multiplying × 2) : C_2H_5

The molar mass of the empirical formula $C_2H_5 = 5 + (2 \times 12) = 29$ g/mol

$$n = \frac{58}{29} = 2$$

The molecular formula of the hydrocarbon : $(C_2H_5)_2 \longrightarrow C_4H_{10}$

$\therefore$ x = 4 and y = 10

Concentration of solutions

The concentration of solutions is expressed in several ways, including the molar concentration **(Molarity)** which indicates the number of moles of solute per liter of solution.

$$\text{Molarity (Molar concentration) (M)} = \frac{\text{Number of moles of solute mol (n)}}{\text{Volume of solution (V) in liters}}$$

Molarity is estimated in mol/L (or M) units.

Example

Calculate the molarity of the solution resulting from dissolving 2 g of sodium hydroxide in distilled water to make a solution its volume is 500 mL

[Na = 23 , O = 16 , H = 1]

Answer

Molar mass of NaOH = 23 + 16 + 1 = 40 g/mol

$$\text{Number of moles NaOH} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{2}{40} = 0.05 \text{ mol}$$

$$\text{Molarity} = \frac{\text{Number of moles of solute}}{\text{Volume of solution in liters}}$$

$$= \frac{0.05}{0.5} = 0.1 \text{ mol/L } (0.1 \text{ M})$$

Dilution of solutions

When a solution of known concentration is diluted with water, the volume of this solution changes while the amount of the solute remains constant, thus the concentration of the solution decreases.

The volume of water needed to be added to a certain volume of a solution of known concentration, in order to change this concentration to a certain amount, can be calculated using the following relation:

Concentration of the concentrated solution M_1 × Its volume V_1 = Concentration of the diluted solution M_2 × Its volume V_2

«Before dilution» «After dilution»

Example

Calculate the volume of water in liters should be added to 1.5 M sodium carbonate solution to obtain one liter of a solution in which each 10 mL contain 0.53 g of solute.

[Na_2CO_3 = 106 g/mol]

Answer

∵ Each 10 mL (0.01 L) $\xrightarrow{\text{Contain}}$ 0.53 g of solute

1000 mL (1 L) $\longrightarrow$? g of solute

$$\text{Mass of the solute in one liter of it} = \frac{0.53 \times 1}{0.01} = 53 \text{ g}$$

$$\text{Number of moles of the solute} = \frac{53}{106} = 0.5 \text{ mol}$$

The concentration of the dilute solution equals its number of moles (0.5 mol) because its volume is (1 L)

$$M_1 \times V_1 = M_2 \times V_2$$

$$1.5 \times V_1 = 0.5 \times 1$$

$$V_1 = \frac{0.5}{1.5} = 0.33 \text{ L}$$

The volume of water should be added to the concentrated solution = 1 – 0.33 = 0.67 L

Assessment on Chapter One

1 Choose the correct answer from the following options:

(1) Which of the following has the smallest molar mass?

[H = 1, O = 16, C = 12, Cl = 35.5]

ⓐ Water ⓑ Carbon Dioxide

ⓒ Methane ⓓ Hydrogen Chloride

(2) How many ions are present in 16 g of anhydrous copper sulphate $CuSO_4$?

[Cu = 63.5 , S = 32 , O = 16]

ⓐ 1.2×10^{23} ⓑ 3.61×10^{23} ⓒ 1.2×10^{24} ⓓ 3.61×10^{24}

(3) What is the mass of 0.2 mol of carbon? **[C = 12]**

ⓐ 2.4×10^{-3} g ⓑ 4.8×10^{-3} g

ⓒ 0.2 g ⓓ 2.4 g

(4) Each of the following represents 0.25 mol of nitrogen gas (at rtp), <u>except</u> ....

[N = 14]

ⓐ 0.5 mol of atoms ⓑ 3.01×10^{23} atom

ⓒ 7 g ⓓ 5.6 L

(5) 264 g of strontium Sr reacts completely with 213 g of chlorine to form a compound with the molecular formula .... **[Sr = 87.6, Cl = 35.5]**

ⓐ SrCl ⓑ $SrCl_2$ ⓒ $SrCl_3$ ⓓ Sr_2Cl

(6) From the reaction: $2NH_{3(g)} + NaClO_{(s)} \longrightarrow N_2H_{4(g)} + NaCl_{(aq)} + H_2O_{(\ell)}$

What is the number of moles and mass of N_2H_4 produced from the reaction of 34 g of NH_3? **[N = 14, H = 1]**

ⓐ 1 mol/ 18 g ⓑ 2 mol/ 36 g

ⓒ 2 mol/ 64 g ⓓ 1 mol/ 32 g

(7) A sample of oxygen gas whose mass is 8 g contains the same number of atoms found in 16 g of element X.

What is the molar mass of element X? **[O = 16]**

ⓐ 4 ⓑ 8 ⓒ 16 ⓓ 32

(8) What is the volume of gas (at STP) produced when 4.6 g of sodium reacts with an excess of water, according to the equation:

$$2Na_{(s)} + 2H_2O_{(\ell)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$$

ⓐ 1.12 L ⓑ 2.24 L ⓒ 4.48 L ⓓ 11.2 L

(9) What is the number of moles of sodium hydroxide in a solution of volume 2 L and concentration 5 M?

ⓐ 2.5 mol ⓑ 5 mol ⓒ 7 mol ⓓ 10 mol

(10) What is the concentration of sodium carbonate solution resulting from adding 240 cm^3 of water to 10 cm^3 of a solution with a concentration of 0.8 M?

ⓐ 0.016 M ⓑ 0.032 M ⓒ 0.04 M ⓓ 0.064 M

---

2 **Calculate the mass of each of:**

(1) 0.2 mol of calcium. **[Ca = 40]**

(2) 0.25 mol of oxygen gas. **[O = 16]**

(3) 1 mol of CO_2 **[C = 12, O = 16]**

(4) 0.5 mol of NaOH **[Na = 23, O = 16, H = 1]**

---

3 **Calculate the number** of molecules in 4.4 g of carbon dioxide gas. **[C = 12, O = 16]**

---

4 **Calculate the number** of atoms in 4 g of hydrogen gas. **[H = 1]**

---

5 **Calculate the number** of ions produced from dissolving 10.6 g of sodium carbonate in water. **[Na = 23 , C = 12 , O = 16]**

---

6 **Calculate the mass** of P_4O_{10} that can be obtained from adding 1.33 g of P_4 to 5 g of O_2 under suitable reaction conditions. **[P = 31, O = 16]**

7 **Calculate the number** of moles of H_2SO_4 that can be prepared from 3 g of the compound Cu_2S, if each sulphur mole turns into a mole of H_2SO_4 **[Cu = 63.5, S = 32]**

8 **Calculate the mass** of NH_4Cl needed to prepare a 100 mL aqueous solution, where each 1 mL contains 70 mg of solute. **[N = 14, H = 1, Cl = 35.5]**

9 **From the reaction:** $C_2H_{4(g)} + 3O_{2(g)} \longrightarrow 2CO_{2(g)} + 2H_2O_{(g)}$

What is the number of moles of the gases produced upon mixing 1 mol of C_2H_4 with 4 mol of O_2 under suitable reaction conditions?

10 **Calculate the volume** of Na_2SO_4 solution with a concentration of 0.3 M required to prepare a solution of 2 L with Na^+ ions concentration of 0.4 M

11 **Calculate the volume** of $BaCl_2$ solution containing 3 mol of chloride ions at 0.5 M

12 A sample of an organic compound with a mass of 10.2 g contains carbon, hydrogen, and oxygen, producing 23.1 g of CO_2 and 4.72 g of H_2O when burned in excess oxygen.

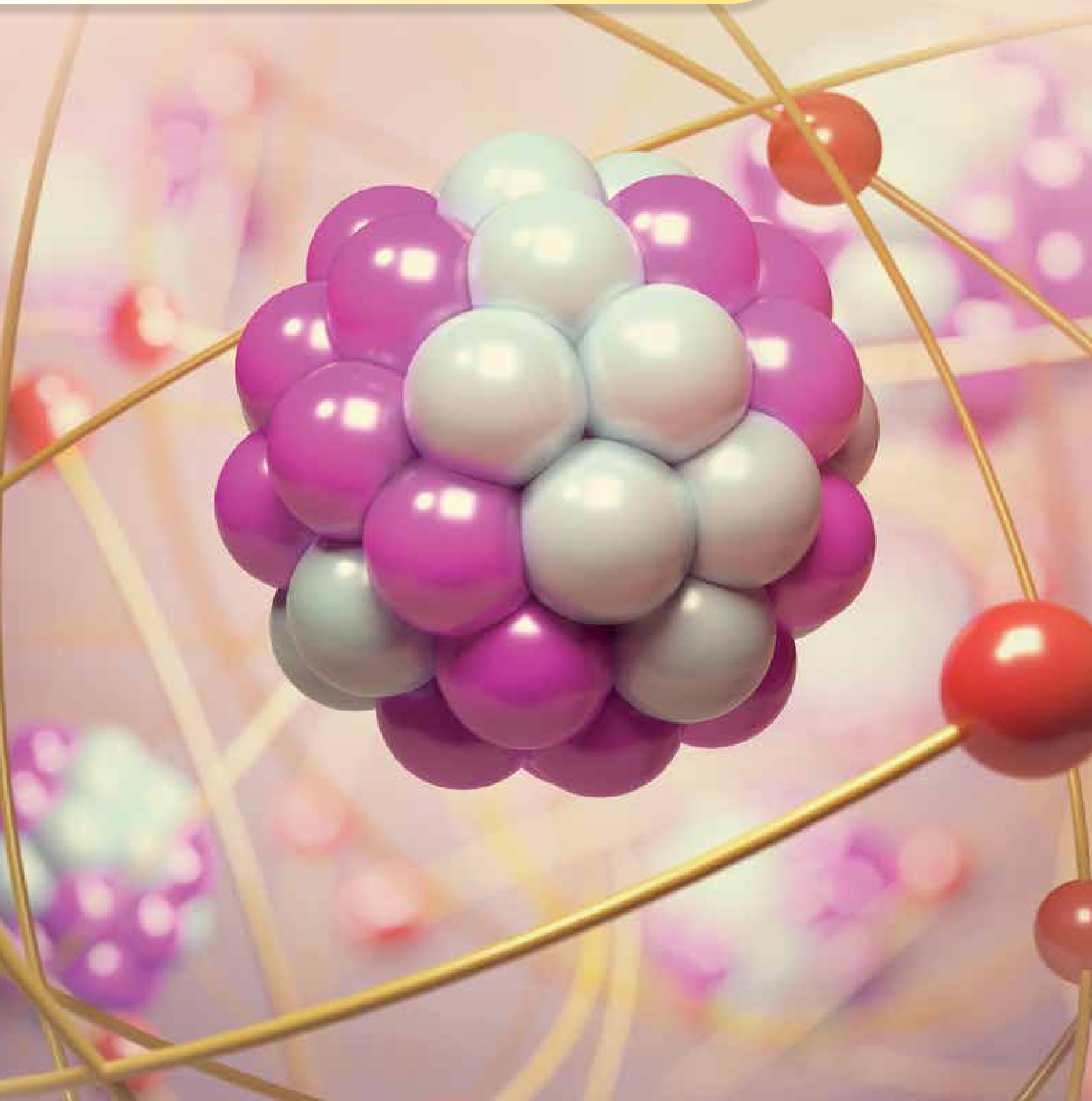
Deduce the empirical formula of the compound mathematically.

[C = 12, H = 1, O = 16]

13 **Calculate the volume** of 2 M hydrochloric acid solution required to prepare 250 cm^3 of a solution its concentration is 0.15 M

Chapter Two

Atomic Structure



Objectives

By the end of this chapter, the student should be able to:

1- Describe Thomson's atomic model.

2- Describe Rutherford's atomic model.

3- Explain the limitations of Bohr's atomic model.

4- Explain key modifications made by modern atomic theory to atomic structure.

5- Define electron cloud and orbital concepts.

6- Identify the four quantum numbers of electrons in atoms.

7- Relate the principal quantum number to sub-levels and orbitals.

8- Apply electron configuration using the Aufbau principle, Hund's rule, and Pauli exclusion principle.

9- Appreciate scientists' contributions to understanding atomic structure.

Atomic Structure

Scientists have made several attempts to comprehend atomic structure, including:

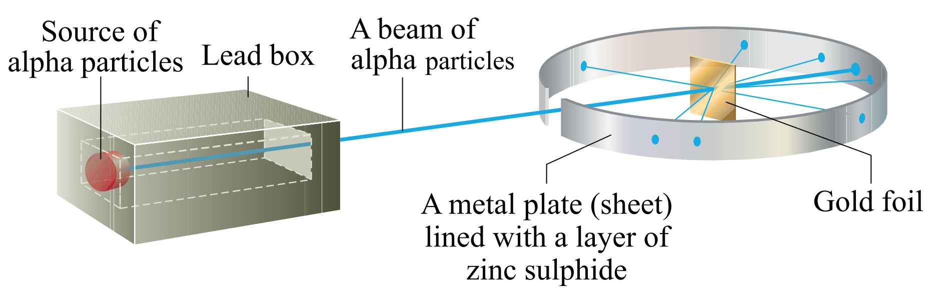
1. Thomson's atomic model.
2. Rutherford's atomic model.
3. Bohr's atomic model.
4. Modern atomic theory principles.

1 Thomson's atomic model

In 1897, Thomson proposed that the atom is a solid (not hollow), homogeneous sphere of uniform positive charge with negative electrons embedded within it, making the atom electrically neutral.

2 Rutherford's atomic model

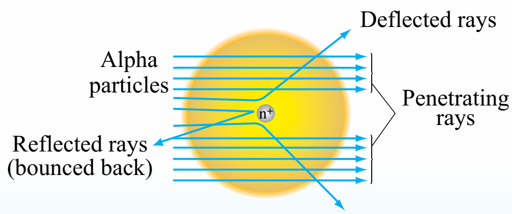
Researchers Geiger and Marsden conducted Rutherford's famous gold foil experiment using the apparatus shown in the following figure:



Rutherford's experiment

Steps of the experiment

1. Rutherford directed a beam of positively charged alpha particles (helium nuclei, ${}^{4}_{2}He$) toward a metal plate coated with zinc sulfide, which produces flashes when struck by alpha particles. Without gold foil present, he determined the positions and number of alpha particles hitting the plate based on flash intensity.
2. Rutherford then placed a very thin gold foil in the path of the alpha particles before they reached the plate. From his observations, Rutherford concluded the following:



The path of alpha particles in a thin gold atom

Observation	Conclusion
1. Most alpha particles showed their effect in the same initial place (penetrated the foil) where they appeared before placing the gold foil.	1. Most of the atom is empty space (**i.e.** it is not a solid sphere, as Thomson persumed).
2. A very small percentage of alpha particles did not pass through the gold foil and bounced back (were reflected) in the opposite direction and some flashes appeared on the other side of the plate.	2. There is a part of the atom having high density and occupies a very small space called the nucleus.
3. Some flashes appeared on both sides of the initial position (were deflected).	3. The charge of the dense part of the atom, where most of its mass is concentrated, must be similar to the positive charge of alpha particles since they repelled each other.

Postulates of Rutherford's atomic model

From the previous experiment and other experiments carried out by other scientists, Rutherford was able to establish the following model (in 1911):

1. The atom :

Despite its extremely small size, it has a complex structure resembling the solar system, which consists of a central nucleus (resembling the sun) around which electrons revolve (resembling planets).

2. The nucleus :

It is much smaller than the atom, and there is a vast space between the nucleus and the electron orbits (meaning the atom is not solid), it is positively charged, and most of the atom mass is concentrated in the nucleus.

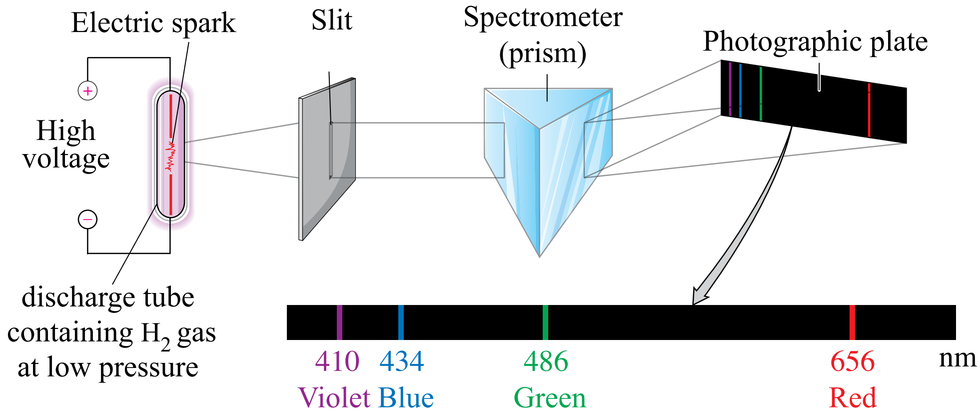
3. Electrons :

1. Their mass is very small (negligible) compared to the mass of the nucleus.
2. The total negative charge of all electrons in the atom equals the positive charge of its nucleus (meaning the atom is electrically neutral).
3. They orbit the nucleus at high speeds in definite orbits, despite the mutual attraction forces between them and the nucleus, the attraction forces are balanced by other forces equal in magnitude but opposite in direction, which are the centrifugal forces.

However, Rutherford's theory did not clarify the system in which electrons orbit the nucleus.

Atomic Emission Spectrum

On heating atoms of a pure element in its gaseous or vapor state to high temperatures, or exposing them to low pressure in an electric discharge tube, radiation is emitted, referred to as the line (emission) spectrum. This spectrum appears when examined by a spectrometer as a limited number of distinct colored lines, each line having a specific wavelength and frequency; these lines are separated by dark areas, hence it is called the line spectrum. It has been experimentally found that the line spectrum of any element is a characteristic property, so no two elements have the same line spectrum.



The emission spectrum appears as distinct coloured lines separated by dark areas

3 Bohr's atomic Model

The study and the explanation of atomic spectra is the key that solved the mystery of atomic structure, which was achieved by the Danish scientist (Niels Bohr) in 1913, for which he was awarded the Nobel Prize in Physics in 1922

Bohr's postulates

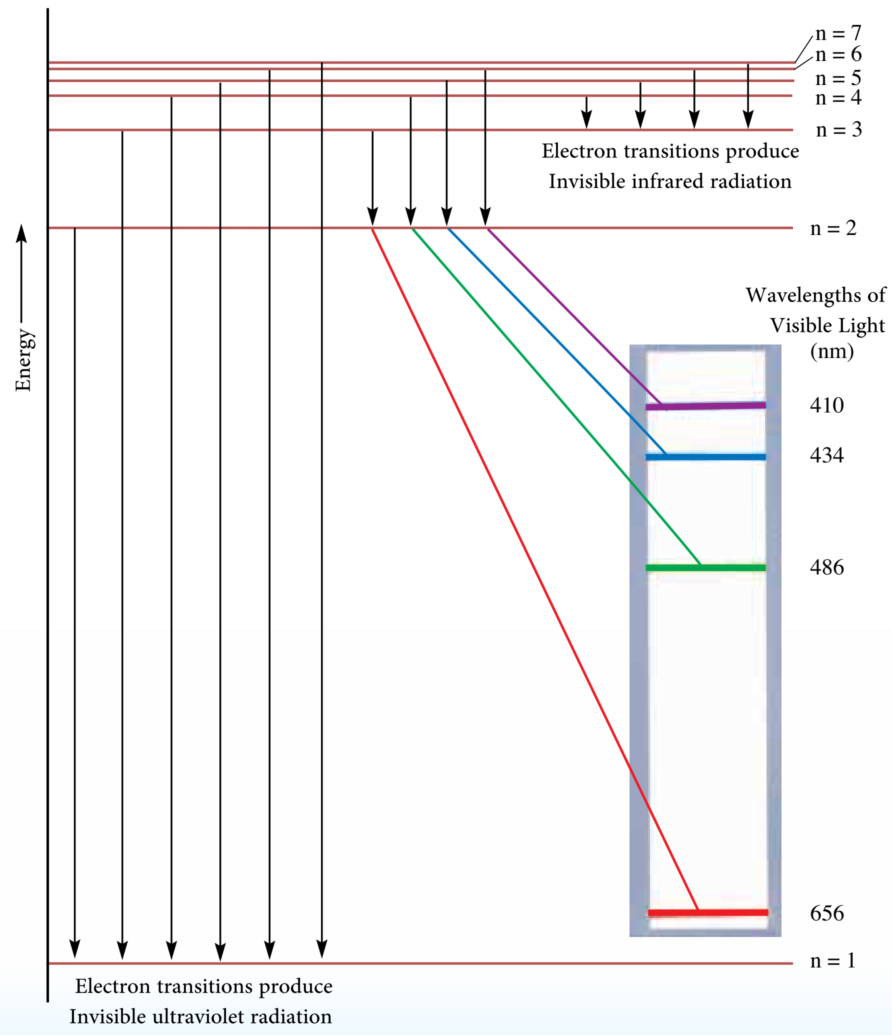
Bohr used some of Rutherford's postulates of the atomic structure, which are:

1. There is a positively charged nucleus at the center of the atom.
2. The number of negative electrons equals the number of positive charges carried by the nucleus.
3. During the electron rotation around the nucleus, a centrifugal force is generated equivalent to the nucleus attraction force to the electron, **then he added the following postulates to Rutherford's postulates**:
4. Electrons move rapidly around the nucleus without losing or gaining any amount of energy.
5. Electrons orbit the nucleus in a number of definite energy levels, and the spaces between these levels are completely forbidden regions for electron rotation.
6. The electron has a certain energy while moving around the nucleus that depends on the distance of its energy level from the nucleus, and the energy of the level increases as its radius increases.

   The energy of each level is expressed by an integer number called the principal quantum number.

7. The electron remains in the lowest possible energy level in a stable state, but if it gains an amount of energy (called a quantum) through heating or electric discharge, the atom becomes excited and the electron temporarily moves to a higher energy level depending on the quantum gained. The electron in the higher level is in an unstable state and soon returns to its original level, losing the same quantum of energy it has gained during its excitation in the form of radiation of light with a distinctive wavelength and frequency, producing a characteristic spectral line.
8. There are many atoms that absorb different quanta (plural of quantum) at the same time that many other atoms emit other quanta. As a result, spectral lines are produced indicating the energy levels between which electrons transfer. This explains the spectral lines of the hydrogen atom.

The lines of visible spectrum of hydrogen appear when the excited electron returns from a higher energy level (except the seventh) to the second energy level, while when the excited electron returns to the first energy level, the limited spectrum is in the ultraviolet region, which is not visible to the naked eye.



The visible line spectrum of hydrogen consists of four coloured lines.

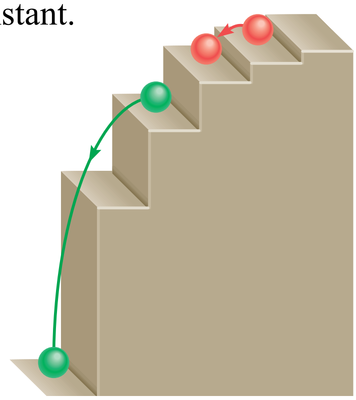
And here are some observations that should be taken into account:

1. Quantum: It is a quantity of energy gained or emitted (released) when an electron transfers from one energy level to another.
2. Bohr's calculations of the radii of the energy levels and the amount of energy of each level showed that the difference in energy between them is not equal, and it decreases as we move away from the nucleus, and thus the quantum of energy required to transfer an electron between different energy levels is not constant.
3. The electron never stabilizes at any distance between energy levels, but moves in definite jumps to the locations of the energy levels.

Bohr's atomic model succeeded greatly in the following:

(a) It correctly explained the hydrogen atom's line spectrum.

(b) Bohr's theory first introduced the quantum concept in determining electron energy at different energy levels.



Electrons jump between any two energy levels without stabilizing in the regions of space between energy levels.

The shortcomings of Bohr's atomic model.

Despite the great efforts made by Bohr to conceptualize the atomic model, the quantitative calculations of his theory did not match many experimental results,

And among the most important shortcomings of Bohr's theory are the following:

1. Bohr's model focused primarily on the hydrogen atom, the simplest electronic system. It successfully explained hydrogen's spectral lines only but failed to explain the spectrum of any other element, even helium, which contains only two electrons.
2. He considered the electron merely a negative material particle and did not account for its wave properties.
3. He assumed that both the position and speed of the electron could be determined precisely simultaneously. In reality, this is practically impossible.
4. Bohr's equations showed the electron as a particle moving in a circular planar system, meaning the hydrogen atom is planar. However, it has been established that atoms are three-dimensional.

These shortcomings prompted scientists to continue their journey toward understanding the true atomic structure.

4 The modern atomic theory

The modern theory of atomic structure is based on fundamental modifications to (Bohr's) model, and the most important of these modifications are as follows:

(A) The dual nature of the electron.

(B) Heisenberg's uncertainty principle.

(C) The wave mechanical theory of the atom.

a The dual nature of the electron:

Previous theories considered the electron merely a tiny negatively charged particle. However, experiments have proven that the electron has a dual nature, **i.e.** it is a material particle that also has wave properties.

b Heisenberg's uncertainty principle:

Bohr's theory assumed that the position and speed of the electron could be determined precisely and simultaneously. Heisenberg concluded an important principle using quantum mechanics. According to this principle, determining both the position and velocity of the electron at the same time is practically impossible. Instead, we can only say that it is probable (to a greater or lesser extent) that the electron exists in a particular location. Speaking in terms of probabilities is a more accurate approach.

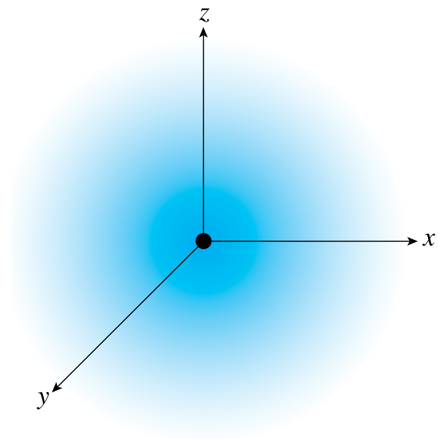
c The wave - mechanical theory of the atom:

In 1926, Austrian scientist Schrödinger established the wave mechanical theory of the atom and formulated the wave equation that can be applied to determine electron motion in the atom. By solving it, we can determine the allowed energy levels and identify regions around the nucleus where the probability of finding electrons in each energy level increases, thus changing our concept of electron motion around the nucleus. Previously, we knew that electrons moved in definite orbits, with regions between these orbits considered forbidden for electrons.

The concept of the electron cloud is used to express the region of space surrounding the nucleus, where the electron is likely to be found from all directions and dimensions .

Within the electron cloud, there are regions where the probability of finding the electron increases, each of which is called orbital.

The mathematical solution to Schrödinger wave equation gave four numbers called quantum numbers.



Electron cloud

Quantum Numbers

To determine the energy of electrons in multi-electron atoms, it is necessary to know the values of four quantum numbers that describe them:

First Principal Quantum Number (n) : describes the electron's distance from the nucleus.

Second Subsidiary Quantum Number (ℓ): describes the shapes of electron clouds in sublevels.

Third Magnetic Quantum Number (m_ℓ): describes orbital shapes and their spatial orientations.

Fourth Spin Quantum Number (m_s) : describes the electron's spin motion around its own axis.

First Principal quantum number (n)

The principal quantum number was previously used by Bohr in explaining the line spectrum of hydrogen atom and is denoted by the symbol (n), and is used to determine the following:

1. The order of the principal energy levels (or electron shells) and their number in the heaviest known atom, in its stable (ground) state, equals seven.

(2) The number of electrons that can fill a certain energy level equals twice the square of the shell number ($2n^2$), where (n) equals the shell number:

The first shell (K) is filled with	$2(1)^2 = 2e^-$
The second shell (L) is filled with	$2(2)^2 = 8e^-$
The third shell (M) is filled with	$2(3)^2 = 18e^-$
The fourth shell (N) is filled with	$2(4)^2 = 32e^-$

This rule does not apply to energy levels higher than the fourth level, as the atom becomes unstable if the number of electrons in any level exceeds 32 electrons.

- The principal quantum number is always a positive integer taking values 1, 2, 3... It does not take the value zero or non-integer values.

Second Subsidiary (secondary) quantum number (ℓ)

- Each principal energy level is divided into a number of sublevels (subshells) with energy determined by a value called the subsidiary quantum number.

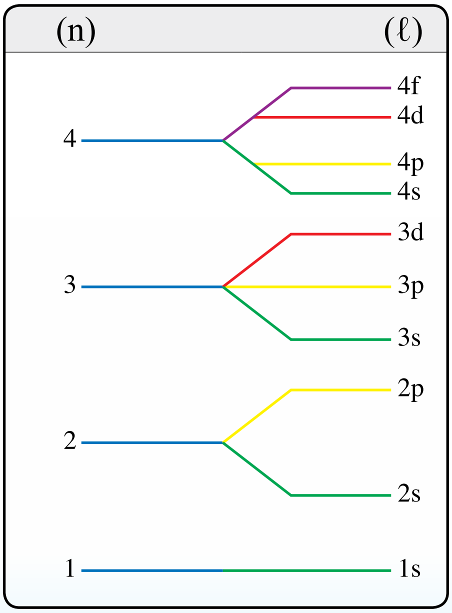
It is characterized by the following:

(A) It determines the energy sublevels in each principal energy level.

(B) The principal level contains a number of energy sublevels equal to its number.

(C) The sublevels take the symbols and values shown in the following table:

Sublevel symbols	s	p	d	f
The values of the subsidiary quantum number (ℓ) [(0) : (n - 1)]	0	1	2	3



It is noted that the sublevels of the same principal level differ slightly in energy and can be arranged according to their energy as follows: f > d > p > s

Third Magnetic quantum number (m_ℓ)

It is characterized by the following:

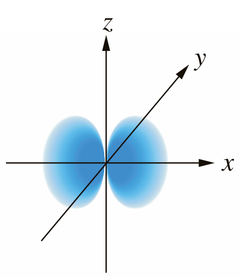
(A) The magnetic quantum number represents the number of orbitals contained in a specific sublevel and their spatial directions.

(B) It is represented by integer values ranging from (-ℓ,...,0,...,+ℓ).

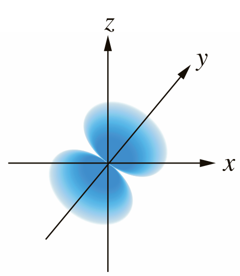
The adjacent table shows the possible values of the magnetic quantum number for level (n=4). The sub-level (s) has one orbital with a spherical shape that is symmetrical around the nucleus, while the sub-level (p) consists of three orbitals that align with the three spatial directions x, y, z, thus denoted by the symbols p_x, p_y, p_z, and they are perpendicular to each other.

The electron density for each of its orbitals takes the shape of two opposite lobes meeting at the center (dumbbell shape) at a point where the electron density is zero.

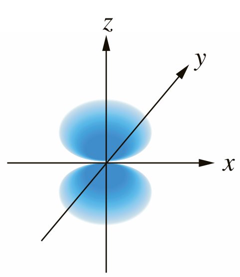
(n)	(ℓ)	(m_ℓ)
1	0	0
2	0	0
	1	– 1,0,+1
3	0	0
	1	–1,0,+1
	2	–2,–1,0,+1,+2
4	0	0
	1	–1,0,+1
	2	–2,–1,0,+1,+2
	3	–3,–2,–1,0,+1,+2,+3



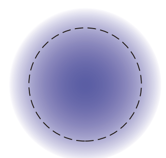
Orbital $2p_x$

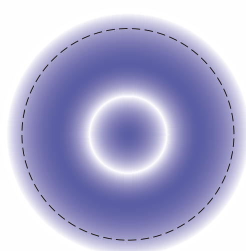


Orbital $2p_y$



Orbital $2p_z$


Orbital **1*s***


Orbital **2*s***

The sublevel (d) consists of five orbitals, while the sublevel (f) consists of seven orbitals.

Fourth Spin quantum number (m_s)

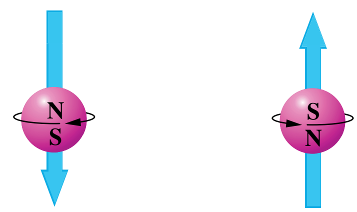
Any orbital can accommodate no more than two electrons. Each electron spins around its own axis while orbiting the nucleus (similar to Earth's rotation around its own axis while revolving around the sun).

Since the two electrons in one orbital carry the same negative charge, one might expect them to repel each other. However, due to each electron's rotation around its axis in a specific direction, a magnetic field is formed that opposes the field generated by the other electron's rotation around its axis. The electrons are said to be in a spin-paired state, denoted by the symbol ($\uparrow\downarrow$).

The following is noted about the spin quantum number:

The spin quantum number (m_s) determines the type of the electron spin motion.

The motion is either clockwise ($\uparrow$) and the value (m_s) is ($+\frac{1}{2}$) or anticlockwise ($\downarrow$) and the value (m_s) is ($-\frac{1}{2}$)



$m_s = +\frac{1}{2}$ $m_s = -\frac{1}{2}$

The spin motion of the two electrons in one orbital

The relationship between the principal quantum number, sublevels, and orbitals:

1. The number of subshells (sublevels) equals the principal quantum number it belongs to, so the first level has one sublevel and the second level has two sublevels .., etc.

2. The number of orbitals in the principal level equals the square of its number n^2, so the second level contains 4 orbitals which are: $2s$, $2p_x$, $2p_y$, $2p_z$ and the third level contains 9 orbitals which are one in the sublevel 3s and three in the sublevel 3p and five in the sublevel 3d

3. The number of orbitals in the sublevel equals $(2\ell + 1)$.

   In the sublevel p whose (ℓ) value is 1, the number of its orbitals is (2 X 1) + 1 = 3 orbitals.

4. The maximum number of electrons that the principal level can hold equals twice the square of its number $2n^2$, so the second level can hold up to 8 electrons distributed as follows:

$$2s^2 , 2p^2_x , 2p^2_y , 2p^2_z$$

The quantum numbers of the electrons up to the third energy level can be summarized as in the following table:

Principal energy level	Principal quantum number (n)	Subsidiary quantum number (ℓ)	Magnetic quantum number (m_ℓ)
	• Determines the principal energy levels	• Determines the energy sublevels. • Number of sublevels = Principal level number	• Determines the number of orbitals of the sublevel.
K	1	1s	$\boxed{\uparrow\downarrow}$
L	2	2s	$\boxed{\uparrow\downarrow}$
		2p	$2p_x$ $2p_y$ $2p_z$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$
M	3	3s	$\boxed{\uparrow\downarrow}$
		3p	$2p_x$ $2p_y$ $2p_z$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$
		3d	$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$

Principles of distributing electrons

2 Pauli's Exclusion Principle

Pauli's principle states that no two electrons in the same atom can have the same four quantum numbers, and the following table illustrates that the two electrons of the sublevel 3s have the same values of the quantum numbers (n, ℓ, m_ℓ),but they have different values for the spin quantum number (m_s) :

The four quantum numbers	n	ℓ	m_ℓ	m_s
The first electron	3	0	0	$+\frac{1}{2}$
The second electron	3	0	0	$-\frac{1}{2}$

3 Aufbau Principle (Building-up Principle)

We already have know that each energy level can contain a number of sublevels that differ slightly in energy, so the actual order of energy in the atom is according to the order of the sublevels.

The Aufbau Principle states that:

Electrons must fill the lower energy sublevels first, then the higher energy sublevels.

The sublevels are arranged in ascending order as follows:

$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d <$ ....

The following is a method for filling the energy sublevels, according to the direction of the arrow.

Energy of 4s

$(n + \ell)$ for 4s

$(n + \ell) = 4 + 0 = 4$

Energy of 3d

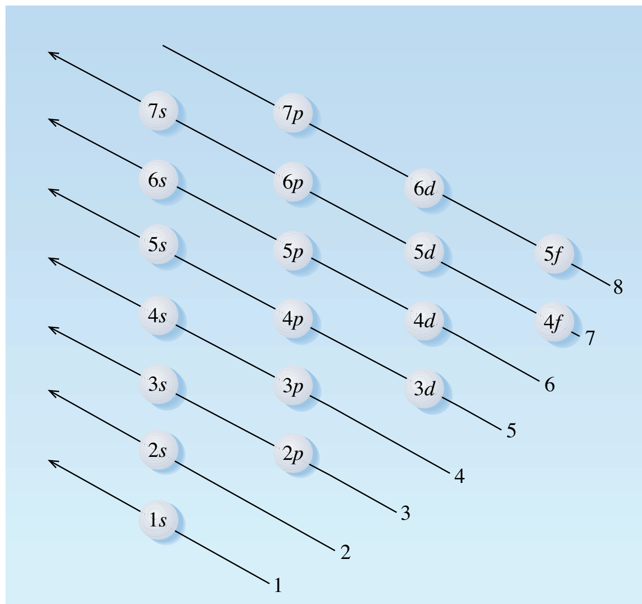
$(n + \ell) = 3 + 2 = 5$

Energy of 4s < Energy of 3d

When the value of $(n + \ell)$ is the same for two sublevels

The sublevel with lower n value has less energy

Energy of 3p < Energy of 4s



The numbers from 1 : 8 represent the sum of $(n + \ell)$ for each energy sublevel

Application

Show the electron distribution of the following atoms according to the Aufbau Principle :

$_{11}Na$, $_{20}Ca$, $_{30}Zn$

Solution

$_{11}Na$: $1s^2$, $2s^2$, $2p^6$,$3s^1$

$_{20}Ca$: $1s^2$, $2s^2$, $2p^6$,$3s^2$, $3p^6$, $4s^2$

$_{30}Zn$: $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$

3 Hund's Rule

Hund's Rule states that no electron pairing occurs in a given sub-level until each orbital contains one electron (filled singly first).

Upon writing the electron configuration of nitrogen (atomic number 7), we find that the sub-level (2p) contains three electrons. Since the sub-level (2p) contains three orbitals that are equal in energy, how are the three electrons distributed among the three orbitals according to Hund's Rule?

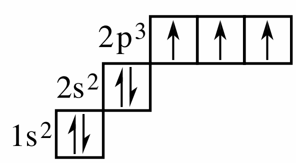
Each electron occupies an independent orbital, as this is energetically favorable. When two electrons pair in one orbital, despite their opposite spins, repulsive forces between them reduce the atom's stability (increasing its energy). Single electrons with parallel spins give the atom maximum possible stability.

For oxygen: The sub-level (2p) has four electrons. Three electrons are distributed first among the three orbitals of the (2p) sub-level according to Hund's rule, while the fourth electron has two possibilities:

- Either it enters one of the three previous orbitals and pairs with an electron already present, where it will experience repulsion with the existing electron.

  Or it can move up to the next sublevel (3s), which has higher energy than (2p). However, it is still energetically favorable for two electrons with opposite spins to pair in the same orbital rather than for one to move to the next higher energy sublevel.

The following example illustrates the electron configuration of nitrogen atom $_7N$ according to Hund's rule:



$$_7N : 1s^2 , 2s^2, 2p^1_x , 2p^1_y , 2p^1_z$$

Assessment on Chapter Two

1 Choose the correct answer from the following options:

(1) The Uncertainty Principle was proposed by ....

ⓐ Schrödinger. ⓑ Rutherford. ⓒ Heisenberg. ⓓ Bohr.

(2) The letters s, p, d, and f to ....

ⓐ principal energy levels. ⓑ energy sublevels.

ⓒ the number of orbitals contained in a sub-level..

ⓓ the number of unpaired electrons in each sublevel.

(3) The quantum number that determines the type of electron movement is the ....

ⓐ principal quantum number. ⓑ subsidiary quantum number.

ⓒ magnetic quantum number. ⓓ spin quantum number.

(4) Which of the following is the electron configuration of nitrogen, according to Hund's rule?

ⓐ 2 , 5 ⓑ $1s^2$, $2s^2$, $2p^3$

ⓒ $1s^2$, $2s^2$, $2p^1_x$, $2p^1_y$, $2p^1_z$ ⓓ $1s^2$, $2s^1$, $2p^4$

(5) If the electron absorbs a quantum of energy, it ....

ⓐ moves to all higher energy levels.

ⓑ moves to the higher energy level that corresponds to the absorbed quantum of energy.

ⓒ moves to any lower energy level.

ⓓ moves to the lower energy level that corresponds to the absorbed quantum of energy.

(6) The magnetic quantum number (m_ℓ) shows...............

ⓐ the number of the principal level in the atom.

ⓑ the number of sublevels.

ⓒ the shape of the orbitals and their spatial directions.

ⓓ the number of electrons in the orbitals.

(7) The number of orbitals in the (3d) sublevel equals ....

ⓐ 5 ⓑ 4 ⓒ 6 ⓓ 7

(8) The number of orbitals in the principal energy level (n) equals ....

ⓐ $2n^2$ ⓑ $3n^2$ ⓒ n^2 ⓓ (n-1)

(9) The maximum number of electrons that can occupy the energy level of the principal quantum number (n) is ....

ⓐ 2n ⓑ n^2 ⓒ $2n^2$ ⓓ $(2n)^2$

(10) The correct arrangement of the following group of energy sublevels according to their energy is ....

ⓐ 3s < 3p< 4d < 4s ⓑ 3s < 4p < 3d < 4f

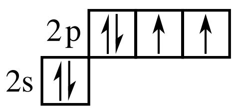
ⓒ 3s < 3p < 3d < 4s ⓓ 3s < 3p < 4s < 3d

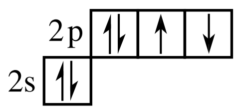
(11) Orbitals of the same energy sublevel are ....

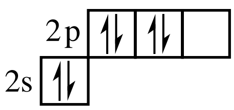
ⓐ different in energy. ⓑ equal in energy.

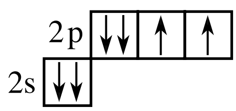
ⓒ different in shape. ⓓ a, c together.

(12) Which of these diagrams shows the correct electron distribution in the last level of the oxygen atom ....

ⓐ 

ⓑ 

ⓒ 

ⓓ 

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2 **Write what explains the following conclusions from Rutherford's experiment observations:**

(1) Most of the atom is empty space and it is not a solid sphere.

(2) There is a part in the atom with a high density that occupies a very small space
(the nucleus of the atom).

(3) The charge of the dense part in the atom, where most of its mass is concentrated, must be similar to the charge of positive alpha particles.

---

3 **Write the possible quantum numbers of the last electron in the following elements:**

(1) Boron $_5B$ **(2)** Fluorine $_9F$ **(3)** Sodium $_{11}Na$

---

4 **Clarify** Thomson's atomic model.

---

5 **What are the possible values of (ℓ)** when (n = 3)?

6 What is meant by each of:

(1) The electron cloud.

(2) The dual nature of the electron.

(3) The aufbau principle.

(4) Hund's rule.

(5) Heisenberg's uncertainty principle.

(6) Pauli's exclusion principle.

---

7 Write the electronic distribution for the following atoms according to the aufbau principle:

(1) $_{35}Br$ **(2)** $_{30}Zn$ **(3)** $_{26}Fe$

(4) $_{10}Ne$ **(5)** $_{20}Ca$ **(6)** $_{11}Na$

---

8 Give reason for the following:

(1) The line spectrum of any element is a characteristic property of it.

(2) The atom is electrically neutral.

(3) Electrons prefer to occupy orbitals singly before pairing in the same sublevel.

(4) The energy sublevel (p) is filled with six electrons, while the energy sublevel (d) is filled with ten electrons.

---

9 Which of the following combinations of quantum numbers contains a mistake? Explain your answer:

(1) $n = 3$, $\ell = 2$, $m_\ell = -1$, $m_s = +\frac{1}{2}$

(2) $n = 4$, $\ell = 3$, $m_\ell = -2$, $m_s = +\frac{1}{2}$

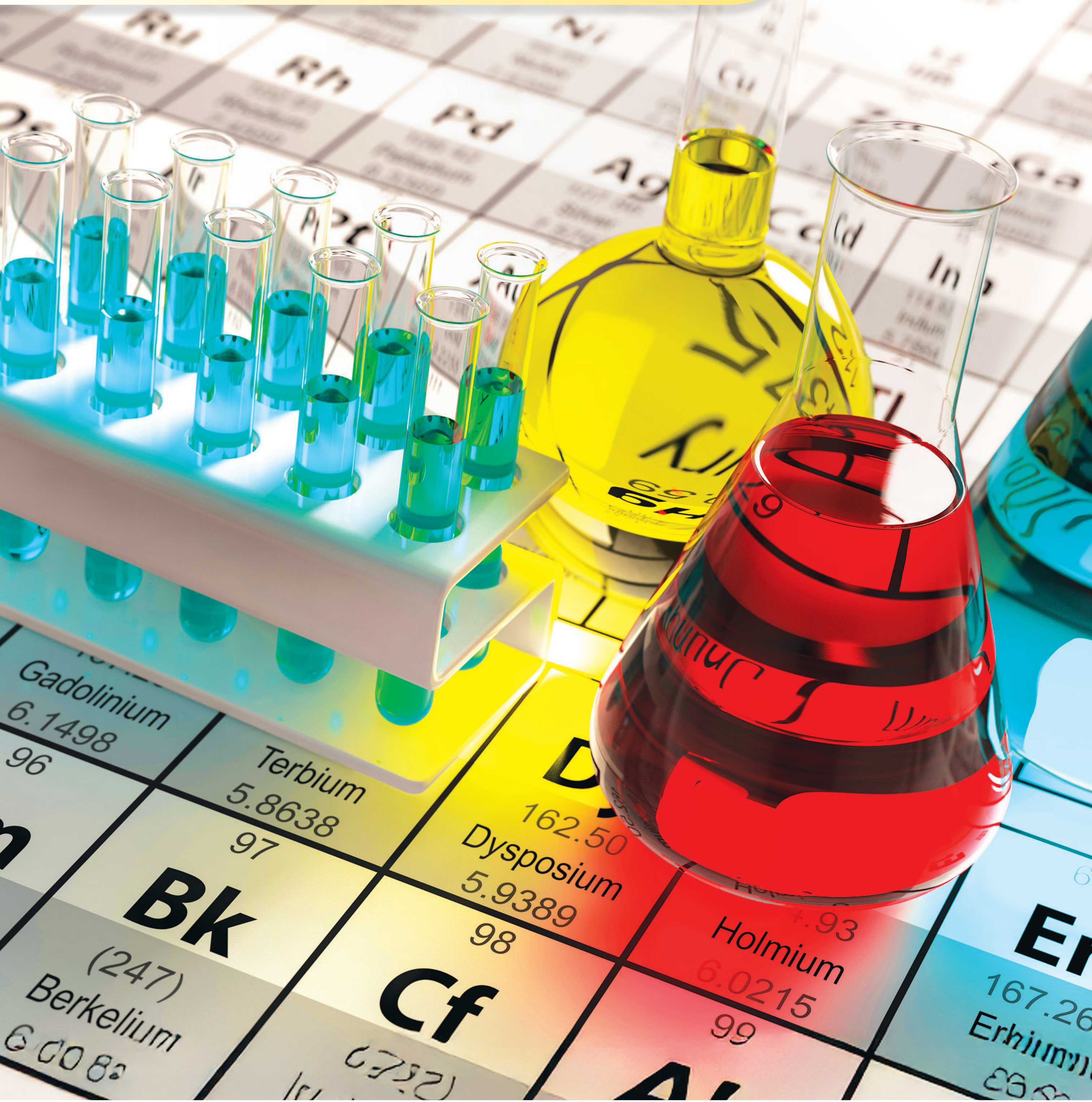
(3) $n = 1$, $\ell = 1$, $m_\ell = +1$, $m_s = -\frac{1}{2}$

---

10 Write the possible values of (ℓ), (m_ℓ) for an electron with a principal quantum number ($n = 2$).

Chapter Three

The Periodic Table and Classification of Elements



Objectives

- By the end of this chapter, the student should be able to:

1- Describe the periodic table.

2- Conclude the type of element and its properties based on its position in the periodic table.

3- Explain why the atomic radius decreases as we move from left to right across a period.

4- Identify the name and location of the blocks in the periodic table.

5- Conclude the relationship between the electronic configuration in a group and the group number.

6- Define atomic radius - ionization energy - electron affinity and electronegativity.

7- Compare electron affinity and electronegativity.

8- Identify the location of metals and nonmetals in the periodic table.

9- Conclude the relationship between atomic radius and both ionization energy and electron affinity in metals and nonmetals.

10- Explain the relationship between atomic number and both basic and acidic properties

11- Calculate the oxidation number of the atom in a compound.

12- Identify oxidation-reduction reactions.

13- Represent oxidation and reduction reactions in half-reactions method.

The modern periodic table

With the development of knowledge about the structure of the atom, the atom true energy levels in the atom were discovered, which are called sub-levels (sub-energy levels), and by reaching the Aufbau principle, the elements were arranged so that each element has one more electron than the previous one.

The fundamental principle on which elements are arranged in the modern periodic table is:

1. Elements are arranged in an ascending order according to their atomic numbers.
2. Elements are arranged according to the Aufbau principle.

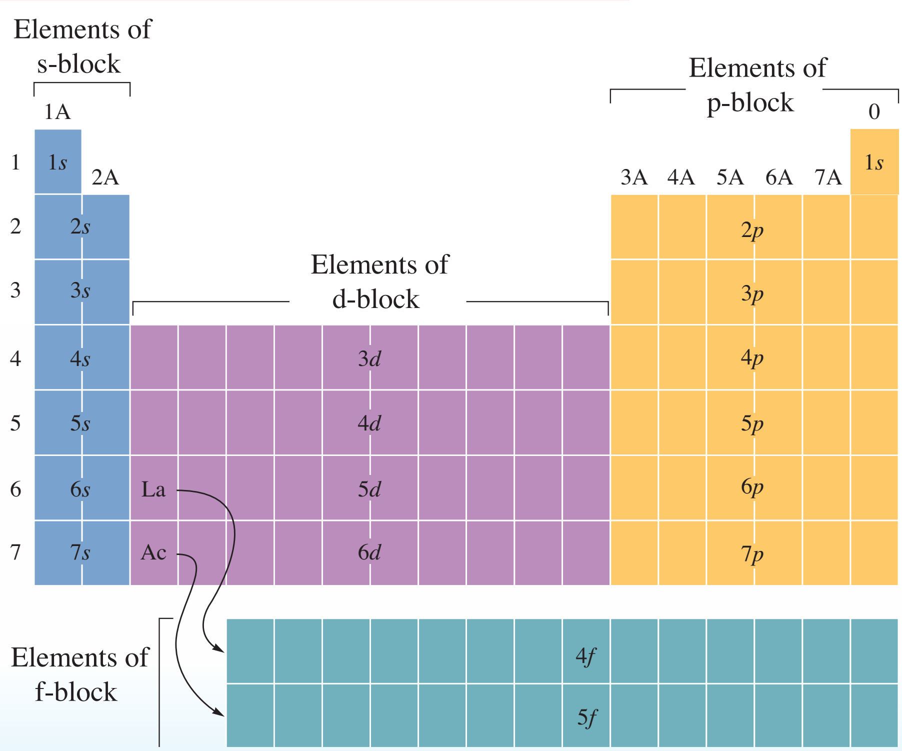
By retrieving the arrangement of sub-levels according to the increase in energy, we find that it agrees with the arrangement of elements in the modern periodic table as follows:

Arrangement of sublevels

1s | 2s < 2p | 3s < 3p | 4s < 3d < 4p

5s < 4d < 5p | 6s < 4f < 5d < 6p | 7s < 5f < 6d < 7p

Blocks of elements in the modern periodic table



Blocks of the modern periodic table

The table is divided into four main areas or blocks

1 Elements of (s) block:

It occupies the left side of the periodic table and contains the elements whose outermost electrons occupy the sublevel (s), which are the elements of group (1A) with the electronic configuration (ns^1) and group (2A) with the electronic configuration (ns^2), where (n) is the last energy level number and the period number at the same time.

2 Elements of (p) block:

It occupies the right side of the periodic table and contains the elements whose outermost electrons occupy the sublevel (p), which are the elements of groups (3A), (4A), (5A), (6A), (7A) and the zero group (noble gases) and the configuration of (p) block elements is (np^1) in group (3A) , (np^2) in group (4A) then the filling of the sublevel (p) continues until it becomes saturated in the zero group (np^6).

3 Elements of (d) block:

They occupy the middle of the periodic table, containing the elements in which the sublevel (d) is filled successively with electrons, and since it can hold up to ten electrons, we find that (d) block consists of ten vertical columns, seven of them belong to the groups (B) and three columns belong to the eighth group.

4 Elements of (f) block:

They are separated down the periodic table so that it does not become too long, which confirms the possibility of separating the periodic table elements into blocks, and the sublevel (f) that can hold up to 14 electrons is filled successively with electrons, this block consists of two series called lanthanides and actinides:

(1) Lanthanides:

In which the sublevel (4f) is filled successively, thus it consists of 14 elements, and it is noted that the outermost valence sublevel of all these elements is ($6s^2$).

(2) Actinides:

In which the sublevel (5f) is filled successively, thus it consists of 14 elements, all of which are radioactive elements and their nuclei are not stable.

Based on what you have learnt, it is clear that the elements of the modern periodic table can be categorized into four types, which are:

1 Noble elements

They are the elements of the last vertical column of the (p) block (group zero or group 18) and their electronic configuration ends by (np^6), except for helium $1s^2$, they are characterized by the filling of all sublevels of the period with electrons, thus they are completely stable elements and form compounds with extreme difficulty.

2 The representative elements

They are the elements of the (s) block and the elements of the (p) block, excluding the elements of group zero. The representative elements are characterized by the filling of all energy levels with electrons except for the last energy level, and they tend to reach the configuration (ns^2, np^6) of their outermost levels by losing or gaining electrons or by sharing.

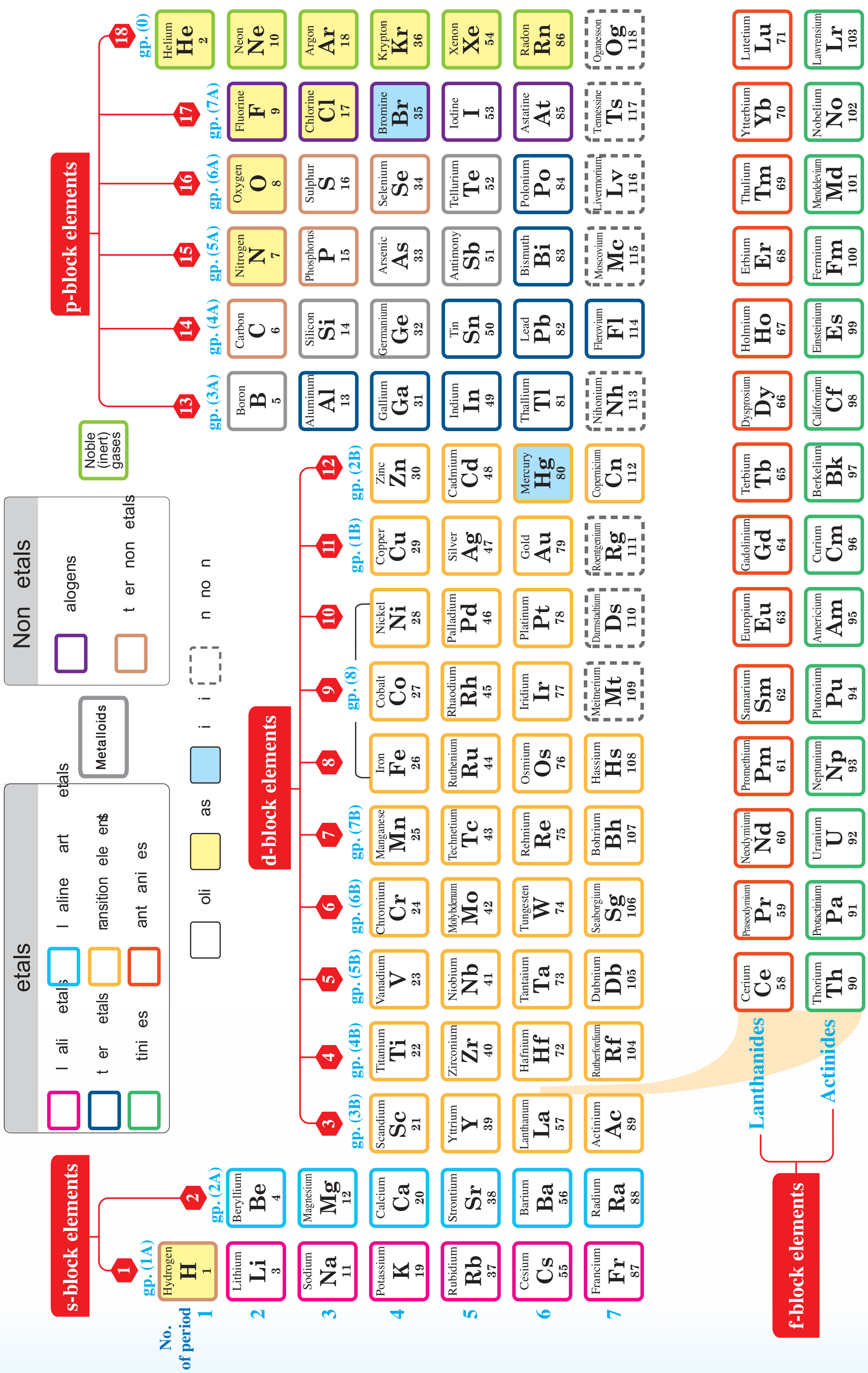
3 The main transition elements

They are the elements of the (d) block, in which the sublevel (d) is filled successively.

4 The inner transition elements

They are the elements of the (f) block in which the sublevel (f) is filled successively.

Elements of the modern periodic table



Description of the periodic table

- The periodic table consists of 18 vertical groups and seven horizontal periods.
- Elements arranged in an ascending order according to the increase in atomic number.
- Each element exceeds the element preceding it in the same period by one electron.
- Each period starts with filling a new energy level with one electron, then filling the sublevels in the same period successively to the last element, which is the noble gas.
- In the vertical groups, elements of the same group are similar in the electron configuration of the last energy level, but with different principal quantum number (n).

The following table shows the electron configuration of some elements arranged according to the atomic numbers up to the element zinc $_{30}Zn$:

The electronic distribution of the first elements in the periodic table											
		1	2		3			4	5	6	7
z	The element	s	s	p	s	p	d	s p d f	s p d f	s p d f	s
1	H	1									
2	He	2									
3	Li	2	1								
4	Be	2	2								
5	B	2	2	1							
6	C	2	2	2							
7	N	2	2	3							
8	O	2	2	4							
9	F	2	2	5							
10	Ne	2	2	6							
11	Na	2	2	6	1						
12	Mg	2	2	6	2						
13	Al	2	2	6	2	1					
14	Si	2	2	6	2	2					
15	P	2	2	6	2	3					
16	S	2	2	6	2	4					
17	Cl	2	2	6	2	5					
18	Ar	2	2	6	2	6					
19	K	2	2	6	2	6		1			
20	Ca	2	2	6	2	6		2			
21	Sc	2	2	6	2	6	1	2			
22	Ti	2	2	6	2	6	2	2			
23	V	2	2	6	2	6	3	2			
24	Cr	2	2	6	2	6	5	1			
25	Mn	2	2	6	2	6	5	2			
26	Fe	2	2	6	2	6	6	2			
27	Co	2	2	6	2	6	7	2			
28	Ni	2	2	6	2	6	8	2			
29	Cu	2	2	6	2	6	10	1			
30	Zn	2	2	6	2	6	10	2			

Trends of properties in the periodic table

After studying the arrangement of elements in the modern periodic table, understanding the electron configuration of elements in the modern table and the relationship between the electron configuration of the element and its position in the table, we will review the trends of physical and chemical properties in the horizontal periods and in the vertical groups and the relationship of these properties to the electron configuration of the elements.

We will focus in this study on the trends of properties in the representative elements, which are the elements of the two blocks (p), (s) and we will leave the study of the trends of properties of transition elements to be researched in another field.

1 Atomic Radius:

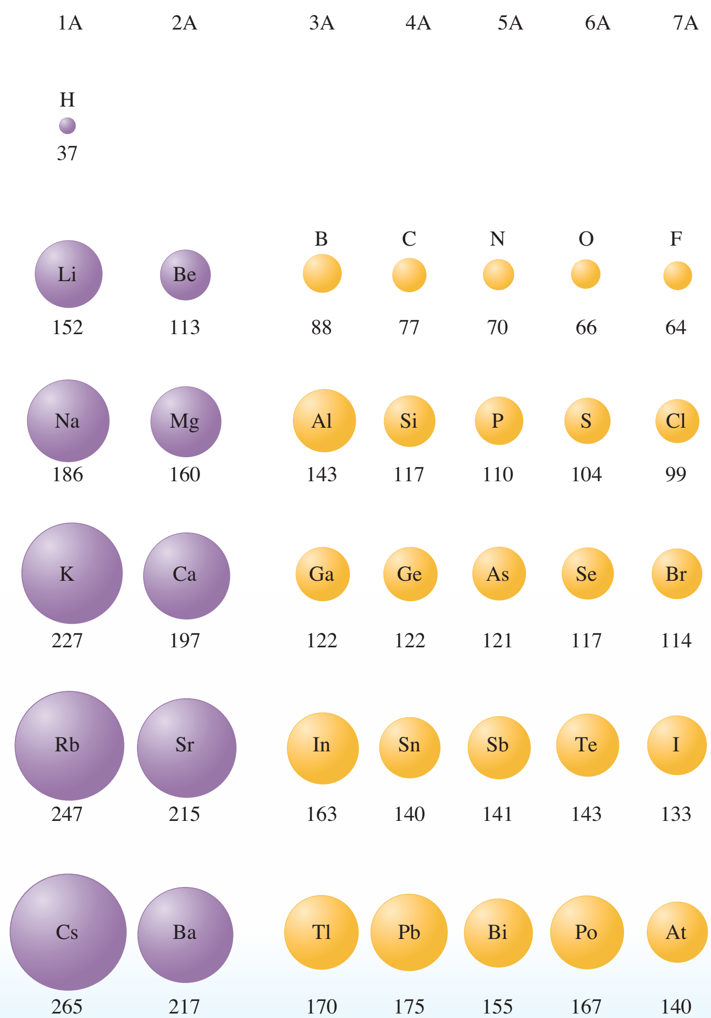
The wave theory has shown that it is impossible to determine the exact location of the electron around the nucleus. So, it is incorrect to define the atomic radius as the distance from the nucleus to the farthest electron, instead **the atomic radius is defined as :**

Atomic radius

- It is estimated by half the distance between the centers of two identical atoms in a diatomic molecule (covalent radius).

It is measured in picometers (pm) and it equals 10^{-12} m. In ionic crystals that consist of positive ions (cations) and negative ions (anions) like sodium chloride crystal, the radius of each is described by the ionic radius. The ionic radius depends on the number of electrons lost or gained.

The opposite figure illustrates the trend of covalent radii of the representative elements in the periodic table:



The trend of atomic radius in picometers (pm)

If we review the radii in the previous figure, we notice the following :

• In the horizontal periods :

The atomic radii decrease as we move to the right from group (1A) to group (7A) due to the gradual increase of the effective nuclear charge (Z_{eff}), which is defined as the actual charge of the nucleus that affects an electron in an atom.

The effective charge is always less than the nuclear charge (number of protons) due to the inner electrons in the complete orbitals (core electrons) shielding part of that charge from the electron under study, thus increasing the attraction of valence electrons, leading to a contraction of

the atomic radius. This means that the largest atoms in size in a single period are the atoms of group (1A), and the smallest in size are the atoms of group 7A (the halogens).

• In the vertical groups :

The atomic radius increases with the atomic number as we move vertically down the same group. **This is due to the following:**

- Increase in the number of energy levels in the atom.
- Filled energy levels shield the effect of the nucleus on the outermost electrons.
- Increased repulsion forces between electrons and each other.

Application

Explanation of the change in the ionic radii compared to the atomic radii of sodium, chlorine, and iron, as shown in the following table:

The atom or ion	Na	Na^+	Fe	Fe^{2+}	Fe^{3+}	Cl	Cl^-
Radius (pm)	186	102	126	78	64.5	99	181
Number of protons	11	11	26	26	26	17	17
Number of electrons	11	10	26	24	23	17	18

In case of metals like sodium, the radius of the positive ion is less than the atomic radius due to the increased effective nuclear charge in the case of the ion. As the charge of the ion increases, as in iron +2 , iron +3 , the radius of the positive ion is less than the atomic radius, due to the increased number of protons compared to the number of electrons.

And in case of nonmetals like chlorine, the radius of the negative ion increases compared to the atomic radius due to the increased number of electrons relative to the number of protons (Do you know the game of tug of war?).

❷ Ionization Potential (Ionization Energy)

When the atom gains an amount of energy, electrons are excited and move to higher energy levels, but if the amount of energy is relatively large, it ejects the weakest electrons bound to the nucleus of the atom, and the atom becomes a positive ion.

Ionization potential is defined as follows:

Ionization potential

- The amount of energy required to remove one mole of electrons from one mole of single atoms in the gaseous state.

And since it is possible to remove one, two, or three electrons from the atom, there is the first, second, and third ionization potentials.., etc.

First ionization potential: Its result is the formation of an ion with a single positive charge.

$$M_{(g)} \longrightarrow M^{+}_{(g)} + e^{-}$$

Second ionization potential: results in the formation of an ion with two positive charges.

$$M^{+}_{(g)} \longrightarrow M^{2+}_{(g)} + e^{-}$$

The first ionization potential varies in the periodic table as follows:

H 1312							He 23 2
Li 2	Be	B 1	C 1	N 1 2	O 131	F 1 1	Ne 2 1
Na	Mg 3	Al	Si	P 1 12	S 1	Cl 12 1	Ar 1 21
K 1	Ca	Ga	Ge 2	As	Se 1	Br 11	Kr 13 1
Rb 3	Sr	In	Sn	Sb 3	Te	I 1	Xe 11
Cs 3	Ba 3	Tl	Pb 1	Bi 3	Po 12	At	Rn 1 3

Ionization energies are measured in kJ/mol

- In periods : Increases ionization energy values as we move to the right due to the increase in effective nuclear charge and the decrease in atomic radius, which leads to an increase in the nucleus's attraction to valence electrons, requiring more energy to remove them from the atom. Thus, ionization energy is inversely proportional to atomic radius.

- In groups: Ionization energy decreases vertically within a group with increasing atomic number because as the number of electron shells increases, the atomic radius increases, and thus the nucleus's attraction to valence electrons decreases, reducing the energy needed to remove them.

- It is observed that:

• The first ionization potential of noble gases (in group zero) is very high due to the stability of their electronic configuration, as it is difficult to remove an electron from a saturated energy level.

- The second ionization potential is greater than the first ionization potential due to the increase in effective nuclear charge, and the third ionization energy for elements in group 2A increases significantly as it causes the breaking of a saturated energy level with electrons, as shown by the ionization energies of magnesium $_{12}Mg$:

$$Mg_{(g)} \longrightarrow Mg^{+}_{(g)} + e^{-} , \quad \Delta H = +737 \quad kJ/mol$$

$$Mg^{+}_{(g)} \longrightarrow Mg^{2+}_{(g)} + e^{-} , \quad \Delta H = +1450 \quad kJ/mol$$

$$Mg^{2+}_{(g)} \longrightarrow Mg^{3+}_{(g)} + e^{-} , \quad \Delta H = +7730 \quad kJ/mol$$

3 Electron Affinity

We learned that the loss of one mole of gas atoms to one mole of electrons is accompanied by the absorption of an amount of energy, which is known as ionization potential.

However when one mole of gas atoms gains one mole of electrons, it either releases or absorbs an amount of energy, which is known as electron affinity.

Electron affinity

The amount of energy changed (released or absorbed) when one mole of single gaseous atoms gains one mole of electrons.

$$Br_{(g)} + e^{-} \longrightarrow Br^{-}_{(g)} \quad \Delta H = -325 \ kJ/mol$$

It is clear from the values of electron affinity shown in the following table:

1A	2A	3A	4A	5A	6A	7A	0
H –72.6							He 0
Li –59.6	Be >0	B –26.7	C –122	N +7	O –141	F –328	Ne +29
Na –52.9	Mg >0	Al –42.5	Si –134	P –72.0	S –200	Cl –349	Ar +35
K –48.4	Ca –2.4	Ga –28.9	Ge –119	As –78.2	Se –195	Br –325	Kr +39
Rb –46.9	Sr –5.0	In –28.9	Sn –107	Sb –103	Te –190	I –295	Xe +41
Cs –45.5	Ba –14	Tl –19.2	Pb –35.2	Bi –91.3	Po –183.3	At –270	Rn +41

Values of electron affinity for elements in blocks s , p

1. Irregularity of the trend of electron affinity in periods and groups , unlike the trends of atomic radius and ionization potential.

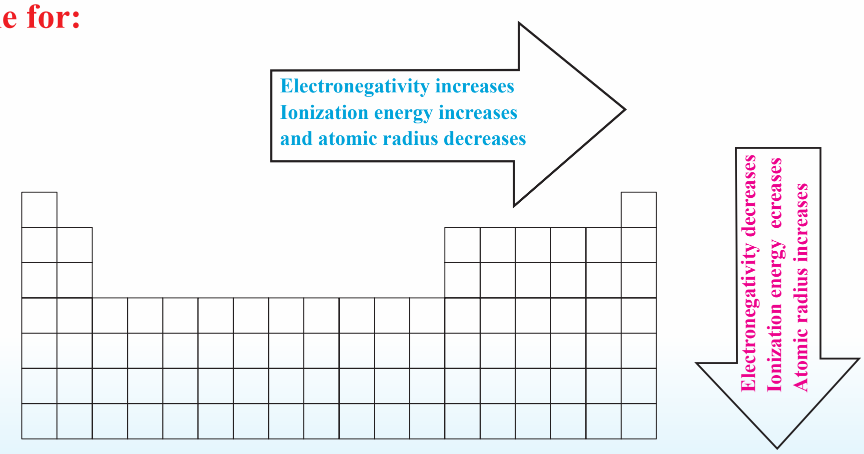
2. The highest energy values of the electron affinity are of the elements of the halogen group (7A), as their anions have same as the electron configuration of the nearest noble gas.
3. The positive electron affinity values for beryllium and magnesium indicate absorption of energy when acquiring electrons, because the added electron occupies the highest energy sublevel (p) compared to the completed sublevel (s).
4. The half-filled 2p sublevel in a nitrogen atom gives it stability, and when it gains an electron, this stability decreases due to pairing of this electron with one of the electrons in the 2p sublevel, which is why the electron affinity value of nitrogen is positive.
5. The electron affinity values of noble gases are positive due to the stability of their electronic configuration.

4 Electronegativity

Electronegativity of an atom is defined as the ability of an atom to attract the electrons of the chemical bond to itself. Electronegativity of element is expressed in value that indicate that an increase in this value corresponds to an increase in its relative ability to attract bonding electrons. We should not confuse electronegativity with electron affinity, as electron affinity is energy related to the atom in its elementary state, while electronegativity refers to the atom when bonded with others.

Electronegativity increases across periods with increasing atomic number and decreasing atomic radius, while in groups, electronegativity decreases with increasing atomic number. Fluorine is considered the highest known element in electronegativity, and the difference in electronegativity between elements plays an important role in determining the type of bonding between them, as will be clarified later (Chapter Four).

The following figure summarizes the gradient of the previous properties in the periodic table for:



5 Metallic and nonmetallic property

The scientist Berzelius was the first to classify elements into two main categories: Metals and nonmetals in the early nineteenth century, and this classification was made before he had any information about atomic structure. Despite the age of this classification, it is still used today, although there are no clear boundaries between the properties of metals and nonmetals. With the development of our understanding of the electronic structure of elements, we can distinguish between metals and nonmetals as follows:

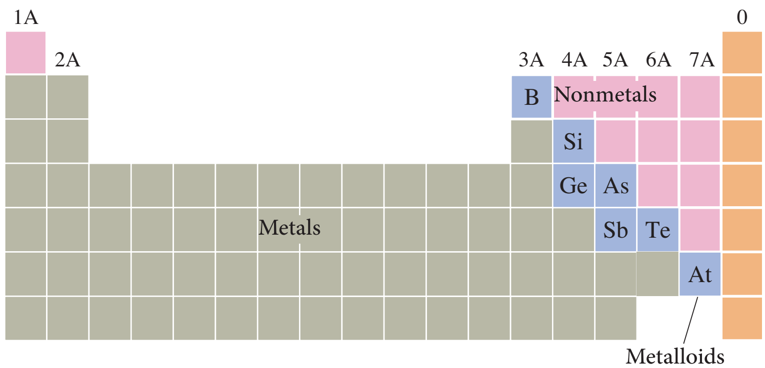
Metals:

1. A group of elements whose valence shell - generally - has less than half of its capacity of electrons.
2. Metals lose valence shell electrons to reach the configuration of the inert gas, which is the goal of chemical reactivity, and become positively charged ions, thus metals are described as electropositive elements.
3. Their ability to conduct electricity is attributed to the ease of transition of their few valence electrons from one place in the metal to another.
4. Metals are characterized by the large radius of their atoms, which leads to a small ionization energy.

Nonmetals:

1. A group of elements whose valence shell - generally - has more than half of its capacity of electrons.
2. They gain a few electrons to reach the noble gas configuration and become negatively charged ions,Thus, nonmetals are described as electronegative elements.
3. Their inability to conduct electricity is attributed to the strong bond of their valence electrons to the nucleus, making it difficult for these electrons to move, and nonmetals are electrical insulators.
4. The small atomic radii of nonmetals lead to a high ionization energy.

There is a third group of elements called Metalloids, which are characterized by having the appearance of metals and most properties of nonmetals. Their electronegativity is intermediate between metals and nonmetals, and their electrical conductivity is less than that of metals but much greater than that of nonmetals. Metalloids are used in the manufacture of parts of electronic devices such as transistors and other semiconductors.



Positionn of metalloids in the periodic table

Reviewing the previous figure, it is clear that all metals are located to the left of metalloids, while nonmetals are located to the right of metalloids. By reviewing the ionization energy and electron affinity of the elements in the periodic table, we find that metallic and non-metallic properties graduated as follows:

- In horizontal periods:

As we move from left to right, we find that the first group contains the strongest metals, then the metallic property gradually decreases with increasing atomic number until we reach metalloids, then the non-metallic property begins to increase until it ends with group 7A, which contains the strongest nonmetals.

- In vertical groups:

We find that the metallic property increases with increasing atomic number as we move down the groups.

From this, we conclude that the strongest metals are located at the bottom left of the periodic table, where cesium is considered the strongest metallic element. The strongest nonmetals are located at the top right of the periodic table, where fluorine is considered the strongest non metallic element..

6 Acidic and basic properties

It is known that nonmetal oxides, when dissolved in water, give acids, such as:

$CO_{2(g)} + H_2O_{(\ell)} \longrightarrow H_2CO_{3(aq)}$ (carbonic acid)

$SO_{3(g)} + H_2O_{(\ell)} \longrightarrow H_2SO_{4(aq)}$ (sulphuric acid)

Non-metal oxides are usually called acidic oxides, and these acidic oxides react with alkalis to produce a salt and water.

$CO_{2(g)} + 2NaOH_{(aq)} \longrightarrow Na_2CO_{3(aq)} + H_2O_{(\ell)}$

Metal oxides are usually called basic oxides, and some basic oxides are soluble in water, while others are not soluble in water. Soluble basic oxides are also called alkaline oxides because they are alkalis, such as:

$Na_2O_{(s)} + H_2O_{(\ell)} \longrightarrow 2NaOH_{(aq)}$

$K_2O_{(s)} + H_2O_{(\ell)} \longrightarrow 2KOH_{(aq)}$

Basic oxides react with acids to produce a salt and water.

$Na_2O_{(s)} + 2HCl_{(aq)} \longrightarrow 2NaCl_{(aq)} + H_2O_{(\ell)}$

$MgO_{(s)} + H_2SO_{4(aq)} \longrightarrow MgSO_{4(aq)} + H_2O_{(\ell)}$

There are two other types of oxides, namely:

- Amphoteric oxides such as aluminum oxide Al_2O_3 and zinc oxide ZnO, and antimony oxide (III) Sb_2O_3 and tin oxide SnO, these oxides react sometimes as basic oxides and other time as acidic oxides.

$Al_2O_{3(s)} + 6HCl_{(aq)} \longrightarrow 2AlCl_{3(aq)} + 3H_2O_{(\ell)}$

$Al_2O_{3(s)} + 2NaOH_{(aq)} \longrightarrow \underset{\text{Sodium aluminate}}{2NaAlO_{2(s)}} + H_2O_{(\ell)}$

$ZnO_{(s)} + H_2SO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + H_2O_{(\ell)}$

$ZnO_{(s)} + 2NaOH_{(aq)} \longrightarrow \underset{\text{Sodium zincate}}{Na_2ZnO_{2(aq)}} + H_2O_{(\ell)}$

- Neutral oxides such as carbon monoxide CO, nitric oxide NO, and nitrous oxide N_2O, which are oxides that react neither with acids nor with alkalis, and do not dissolve easily in water.

Acidic and basic properties in the periodic table

a In horizontal periods

As the atomic number of the element increases, the basic property of the oxide decreases, while the acidic property increases.

b In vertical groups

If we take the elements of group one as an example, we find that the basic property increases vertically as the atomic number increases (**i.e.** as we go down). On the other hand, when tracing the acidic property in the hydrides of the elements of group 7A, we find that as the atomic radius of the element increases, the attraction of hydrogen atom decreases, making it easier to be ionized, thus increasing the acidic property.

7 Oxidation Numbers

In the past, before understanding the nature of chemical bonds, the concept of valency was used to express the ability of an atom to bond with a certain number of other atoms, but now the concept of oxidation number is used as a better alternative to valency.
The oxidation number is defined as:

Oxidation number

- A number that refers to the electric charge (positive or negative) that the atom or ion would carry in the compound, whether it is an ionic or a covalent compound.

To determine the oxidation number of an atom in a compound, the following steps are followed:

First: in ionic compounds:

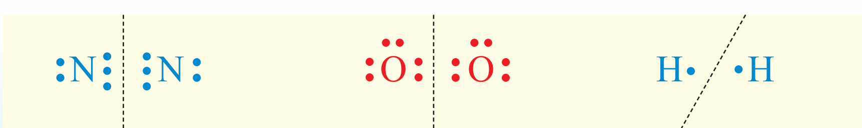
The oxidation number of the ion or atomic group in the ionic compound is equal to its charge and has the same sign, for example:

K^+Br^-, Na^+Cl^- , $Mg^{2+}O^{2-}$, $Ca^{2+}\ CO_3^{\ 2-}$, $Cu^{2+}SO_4^{\ 2-}$

Second: in covalent compounds:

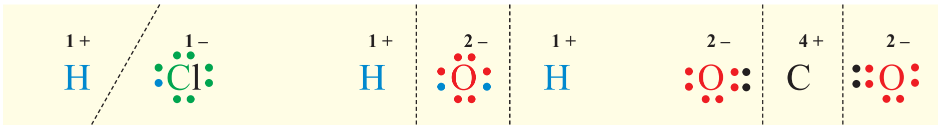
Although the molecule of the covalent compound is not made up of positive and negative ions, its atoms carry charges that indicate the electronic displacement occurring in the bond, where the more electronegative atom carries a negative charge and the less electronegative atom carries a positive charge, and there are two cases:

- In molecules with identical atoms such as S_8, P_4, O_3, Cl_2 the electronic displacement in the bonds between the atoms is equal because the atoms of any molecule of a single element are equal in electronegativity, and thus the oxidation number of any atom in this molecule equals (0).



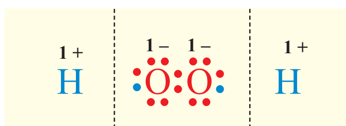
Nitrogen molecule N_2 Oxygen molecule O_2 Hydrogen molecule H_2

- When the molecule consists of two different atoms in electronegativity, the shared electrons are counted withthe atom having more electron negativity noting that the oxidation number of oxygen in most of its compounds is (–2) and the oxidation number of hydrogen in most of its compounds is (+1).



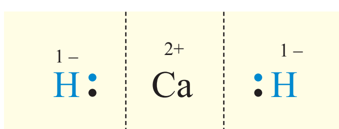
Hydrochloric acid molecule HCl   **Water molecule** H_2O **Carbon dioxide molecule** CO_2

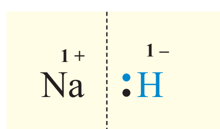
Although the oxidation number of oxygen in most of its compounds is (–2), in peroxides such as hydrogen peroxide (water of oxygen) H_2O_2 and sodium peroxide Na_2O_2 the oxidation number of oxygen is (–1). Moreover, in potassium superoxide (KO_2) the oxidation number of oxygen is ($-\frac{1}{2}$) and in oxygen difluoride (OF_2) the oxidation number of oxygen is (+2).



Hydrogen peroxide H_2O_2

The oxidation number of hydrogen in most of its compounds is (+1), except for some cases such as in the case of hydrides of active metals such as sodium hydride NaH and calcium hydride CaH_2 . These are ionic compounds containing negative hydrogen ion. So, if sodium hydride, for example, is melted for example and electrolyzed, hydrogen is released at the anode - and the oxidation number of hydrogen in the metal hydrides is (–1)





Calcium hydride CaH_2 **Sodium hydride** NaH

When calculating oxidation numbers, the following should be considered:

1. The sum of oxidation numbers for different elements in the neutral molecule is (0) zero.
2. The oxidation number is calculated to one atom or one ion only in the molecule.
3. Some compounds such as $NH_4^+NO_2^-$ have nitrogen atoms with different oxidation numbers.
4. The oxidation number of elements in group (1A) is always (+ 1) and elements of group two (2A) is (+2) and elements of group three three (3A) is (+3). The oxidation number of fluorine is always (–1). So, we start by writing its oxidation number first,, then continue calculating the oxidation numbers of the other elements.

Compound	Na_2O Sodium oxide		Na_2O_2 Sodium peroxide		KO_2 Potassium superoxide		CaH_2 Calcium hydride		AlH_3 Aluminum hydride	
Total charges	2+	2–	2+	2–	1+	1–	2+	2–	3+	3–
Atoms	Na_2	O	Na_2	O_2	K	O_2	Ca	H_2	Al	H_3
Oxidation number	+1	-2	+1	–1	+1	$-\frac{1}{2}$	+2	-1	+3	-1

5. The highest oxidation state of any representative element does not exceed its group number, according to the numbering 1A , 2A , ...

6. The oxidation number for atomic groups = the charge carried by the group, such as:

NO_3^-	CO_3^{2-}	SO_4^{2-}	NH_4^+
Nitrate group (-1)	Carbonate group (-2)	Sulphate group (-2)	Ammonium group (+1)

Oxidation numbers are useful in determining the type of change occurring to the element during the chemical reaction, in oxidation-reductionn reactions.

Oxidation

The process of losing electrons resultsing in an increase in the positive charge

Reduction

The process of gaining electrons resulting in a decrease in the positive charge

The reactant that loses electrons is described as the reducing agent, and the reactant that gains these electrons is described as the oxidizing agent.

When the oxidation number of the central atom in the compound is at its maximum, it acts as an oxidizing agent only.

$$K_2\overset{6+}{Cr_2}O_7 , H\overset{5+}{N}O_3 , H_2\overset{6+}{S}O_4$$

However, when the oxidation number of the central atom in the compound is at its minimum, it acts as reducing agent only.

$$\overset{3-}{N}H_3 , \overset{4-}{C}H_4 , \overset{0}{H_2}$$

In cases where the oxidation state of the central atom in the compound is intermediate, it can act as either an oxidizing agent or a reducing agent depending on the reaction.

$$H\overset{3+}{N}O_2 , \overset{4+}{S}O_2$$

Oxidation-Reduction Reactions (Redox Reactions)

In the redox reactions, the number of electrons lost by the reducing agent during its oxidation equals the number of electrons gained by the oxidizing agent during its reduction. Oxidation and reduction processes in chemical reactions can be verified by tracking the changes in oxidation numbers occurring in them.

Application

In the oxidation-reduction reaction: $2NaI_{(aq)} + Cl_{2(g)} \longrightarrow 2NaCl_{(aq)} + I_{2\,(s)}$

Both NaI and NaCl are in ionic form, so the equation can be written in another form known as the overall equation, as follows:

$$2Na^{+}_{(aq)} + 2I^{-}_{(aq)} + Cl_{2(g)} \longrightarrow 2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)} + I_{2(s)}$$

And the ion Na^{+} in this reaction is described as spectator ion because it does not participate actively in the occurring reaction (no change in its oxidation number on both sides of the equation), and when the spectator ion is removed from the overall equation, we obtain net ionic equation which contains the atoms or ions or molecules that have exhibit charging in their oxidation numbers.

$$2I^{-}_{(aq)} + Cl_{2(g)} \longrightarrow 2Cl^{-}_{(aq)} + I_{2(s)}$$

It is observed in this reaction that:

- Negative iodide ions have been converted into a neutral iodine molecule, indicating an increase in the oxidation number from (–1) : (0)
- The neutral chlorine molecules have been turned into negative chloride ions, indicating a decrease in the oxidation number from (0): (–1)
- The increase in oxidation number represents an oxidation process, and the decrease represents a reduction process. The previous net ionic equation can be divided into two equations representing the oxidation and reduction reactions, each individually as follows:

$2I^{-}_{(aq)} \longrightarrow I_{2(s)}$ ، $Cl_{2(g)} \longrightarrow 2Cl^{-}_{(aq)}$

To achieve the law of conservation of charges, the lost and gained electrons are added to the two half reactions, as follows:

• Half-reaction equation for oxidation: $2I^{-}_{(aq)} \longrightarrow I_{2(s)} + 2e^{-}$

• Half-reaction equation for reduction: $Cl_{2(g)} + 2e^{-} \longrightarrow 2Cl^{-}_{(aq)}$

Balancing ionic equations for the redox reactions:

• **Balancing is done by following these steps:**

Step (1): Separating the oxidation and reduction reactions.

Step (2): Balancing the atoms and ions on both sides of each equation.

Step (3): Applying the law of conservation of charge by adding the lost and gained electrons.

Step (4): Equalizing the number of lost and gained electrons in the oxidation and reduction equations

Step (5): Adding the oxidation and reduction equations.

Application

Balancing the equation: $\mathbf{Cr_{(s)} + H^+_{(aq)} \longrightarrow Cr^{3+}_{(aq)} + H_{2(g)}}$

The step (1):	$Cr_{(s)} \longrightarrow Cr^{3+}_{(aq)}$	$H^+_{(aq)} \longrightarrow H_{2(g)}$
The step (2):	$Cr_{(s)} \longrightarrow Cr^{3+}_{(aq)}$	$2H^+_{(aq)} \longrightarrow H_{2(g)}$
The step (3):	$Cr_{(s)} \longrightarrow Cr^{3+}_{(aq)} + 3e^-$	$2H^+_{(aq)} + 2e^- \longrightarrow H_{2(g)}$
The step (4):	$2Cr_{(s)} \longrightarrow 2Cr^{3+}_{(aq)} + 6e^-$	$6H^+_{(aq)} + 6e^- \longrightarrow 3H_{2(g)}$

The step (5): $2Cr_{(s)} + 6H^+_{(aq)} \longrightarrow 2Cr^{3+}_{(aq)} + 3H_{2(g)}$

- There is a type of redox reactions, where oxidation and reduction processes occur for one element in one reactant, and such reactions are known as disproportionation reactions.

Oxidation

Reduction

$$2H_2\overset{1-}{O}_{2(aq)} \longrightarrow 2H_2\overset{2-}{O}_{(\ell)} + \overset{0}{O}_{2(g)}$$

- And there is another type known as comproportionation reactions. (Symporportionation Reactions)

It occurs for two compounds or ions that involve a common element in two different oxidation states (which combine together) to form a product with an intermediate oxidation state.

Oxidation

$$2H_2\overset{2-}{S}_{(g)} + \overset{4+}{S}O_{2(g)} \longrightarrow 3\overset{0}{S}_{(g)} + 2H_2O_{(\ell)}$$

Reduction

Assessment on Chapter Three

1 Choose the correct answer from the following options:

(1) How many types of elements are present in the sixth period of the modern periodic table?

ⓐ 3 ⓑ 4 ⓒ 5 ⓓ 6

(2) Nonmetals are characterized by ....

ⓐ their ionization energy is high.
ⓑ their elements are electropositive.
ⓒ their electronegativity is low.
ⓓ their atomic radius is large.

(3) Electronegativity increases in horizontal periods ....

ⓐ with increasing atomic radius.
ⓑ with decreasing atomic number.
ⓒ with decreasing atomic radius..
ⓓ (a , b) together.

(4) The following reactions represent oxidation and reduction reactions, <u>except</u> ....

ⓐ $2P + 5HClO + 3H_2O \longrightarrow 2H_3PO_4 + 5HCl$

ⓑ $Zn + 2HCl \longrightarrow ZnCl_2 + H_2$

ⓒ $Mg + CuSO_4 \longrightarrow MgSO_4 + Cu$

ⓓ $NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$

(5) Which of the following reactions represents an oxidation-reduction reaction?

ⓐ $CuO + H_2SO_4 \longrightarrow CuSO_4 + H_2O$

ⓑ $CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$

ⓒ $Cr_2O_7^{2-} + 3H_2S + 8H^+ \longrightarrow 2Cr^{3+} + 3S + 7H_2O$

ⓓ $NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$

2 **Choose from column (B) the electron configuration of the last energy level that fits each element in column (A), then identify the type of element from column (C):**

(A) Element	(B) Electron configuration in the outer energy sub level	(C) Type of element
(1) Radon $_{86}Rn$	ⓐ $7s^1$	I- An actinides.
(2) Cesium $_{55}Cs$	ⓑ $4f^{14}, 5d^6, 6s^2$	II- Transition metal in the fifth period.
(3) Bromine $_{35}Br$	ⓒ $6s^2, 5d^{10}, 6p^6$	III- Noble gas.
(4) Vanadium $_{23}V$	ⓓ $3d^3, 4s^2$	IV- Main transition metal in the sixth period.
(5) Osmium $_{76}Os$	ⓔ $4f^7, 5d^1, 6s^2$	V- One of lanthanides.
(6) Gadolinium $_{64}Gd$	ⓕ $4s^2, 3d^{10}, 4p^5$	VI- Representative element in (s) block.
	ⓖ $4d^5, 5s^1$	VII- Transition metal in the fourth period.
	ⓗ $6s^1$	VIII- Representative element in (p) block.

3 **What is meant by:**

(1) Atomic number.
(2) Reduction.
(3) Representative elements.
(4) Noble gases.
(5) Transition elements.
(6) Inner transition elements.
(7) Atomic radius.
(8) Ionization energy.
(9) Electron affinity.
(10) Electronegativity.
(11) Metals.
(12) Nonmetals.
(13) Metalloids.
(14) Acidic oxide.
(15) Basic oxide.
(16) Amphoteric oxide.
(17) Neutral oxide.
(18) Oxidation number.
(19) Oxidation.
(20) Reducing agent.

4 **Write a brief about the trends of the following properties in the periodic table:**

(1) Atomic radius. **(2)** Ionization energy.

(3) Electronegativity. **(4)** Metallic and nonmetallic property.

5 **What is the difference between each of:**

(1) First ionization potential and second ionization potential.

(2) Electron affinity and electronegativity.

(3) Metals and nonmetals.

(4) Acidic oxide, basic oxide and amphoteric oxide.

(5) Oxidation and reduction.

6 **What is the scientific term of the following:**

(1) Half the distance between the centers of two identical atoms in a diatomic molecule.

(2) The amount of energy required to remove the least tightly bound electrons from one mole of isolated atoms in the gaseous state.

(3) The amount of energy released or absorbed when one mole of gaseous atoms gains one mole of electrons.

(4) The ability of an atom to attract the electrons of the chemical bond.

(5) A group of elements whose valence shell - generally - has less than half the capacity of electrons.

(6) A group of elements whose valence shell - generally - has more than half the capacity of electrons.

(7) A number that refers to the electric charge that the atom would carry in the compound.

(8) The process of losing electrons resulting in an increase in positive charge.

(9) The process of gaining electrons resulting in a decrease in positive charge.

7 **Calculate the oxidation numbers of the following elements:**

(1) Oxygen in: OF_2 - KO_2 - Na_2O_2 - Li_2O - O_3 - O_2

(2) Chlorine in: $NaCl$ - $NaClO_4$ - $NaClO_3$ - $NaClO_2$ - $NaClO$

(3) Nitrogen in: HNO_3 - NO_2 - HNO_2 - NO - N_2O - N_2 - NH_3

(4) Sulphur in: $Na_2S_2O_3$ - K_2S - SO_2 - $NaHSO_3$ - H_2SO_4

(5) Manganese in: $NaMnO_4$ - $MnCl_2$ - $KMnO_4$ - MnO_2

8 Indicate the oxidation or reduction processes (if found) in each of the following:

(1) $CO \longrightarrow CO_2$

(2) $Cr_2O_7^{2-} \longrightarrow Cr_2O_3$

(3) $O_2 \longrightarrow O_3$

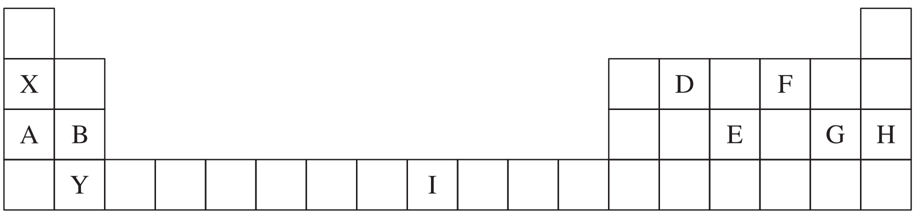
(4) $NO_2 \longrightarrow N_2O_4$

(5) $MnO_4^- \longrightarrow MnO_2$

(6) $ClO^- \longrightarrow ClO_3^-$

(7) $FeCl_3 \longrightarrow FeCl_2$

9 The following figure represents the first four periods of the modern periodic table:



(1) Arrange the following elements according to the decrease in atomic radius B, A, G, E

(2) Arrange the following elements according to the increase in ionization energy F, X, D

(3) What is the block to which each of the elements Y, H, G, I, X belong?

10 Illustrate with balanced symbolic equations that:

(1) Sodium oxide is a basic oxide.

(2) Sulphur trioxide is an acidic oxide.

(3) Zinc oxide is an amphoteric oxide.

11 **The reaction of calcium with water is expressed by the following symbolic equation:**

$$Ca_{(s)} + 2H_2O_{(\ell)} \longrightarrow Ca(OH)_{2(aq)} + 2H_{2(g)}$$

(1) Express the occurring reaction with an ionic equation.

(2) Deduce the balanced two half-reactions equations of both the oxidation and the reduction processes.

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12 **Identify the oxidation, the reduction, the oxidizing agent, and the reducing agent in the following reactions:**

(1) $MnO_{2(s)} + 4HCl_{(aq)} \longrightarrow MnCl_{2(aq)} + Cl_{2(g)} + 2H_2O_{(\ell)}$

(2) $Cl_{2(g)} + H_2O_{(\ell)} \longrightarrow HCl_{(aq)} + HClO_{(aq)}$

(3) $SO_{2(g)} + Cl_{2(g)} + 2H_2O_{(\ell)} \longrightarrow H_2SO_{4(aq)} + 2HCl_{(aq)}$

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13 **Use oxidation numbers to balance the following equations:**

(1) $Cu^{2+}_{(aq)} + Al_{(s)} \longrightarrow Al^{3+}_{(aq)} + Cu_{(s)}$

(2) $Sn^{2+}_{(aq)} + Fe^{3+}_{(aq)} \longrightarrow Fe^{2+}_{(aq)} + Sn^{4+}_{(aq)}$