SLO 3: Atomic Structure

3.1 Features of an Atom

3.1.1 Describe The Structure Of An Atom With Reference To The Location, Relative Electric Charges And Relative Masses Of Proton, Electron And Neutron

The idea of atom was first proposed in Greece when the philosopher Democritus declared that all matter is made of tiny particles. John Dalton further provided evidence and named this particle as atom, a particle that cannot be further subdivided. 'Discharge Tube Experiments' in late nineteenth century discovered protons and electrons. The presence of negatively charged particles was ascertained because of their deflection towards the positive plate in an electric field. Similarly, the presence of positively charged particles was confirmed due to their deflection towards the negative plate.

Thus, atoms are no longer the smallest particles of matter; rather there exist particles that are even smaller than atoms. In other words, atoms are composed of negatively charged particles called electrons and positively charged particles called protons. It was also discovered that a proton is 1836 times heavier than an electron. Later it was found out that neutrons also exist alongside protons inside the nucleus that have almost equal mass as proton but carry no charge.

PARTICLE	LOCATION	CHARGE	MASS
Electron	K,L,M,N shells	-1.6022×10^{-19} C	9.109×10^{-31} Kg
Proton	Inside Nucleus	$+1.6022 \times 10^{-19}$ C	1.673×10^{-27} Kg
Neutron	Inside Nucleus	0 C	1.675×10^{-27} Kg

3.1.2 Define The Following Terms:

A. Atomic (Proton) Number

- The atomic number is the number of protons in the nucleus of an atom.
- It determines the identity of an element.
- It is also equal to the number of electrons in a neutral atom.
- Example: Hydrogen has an atomic number of 1.

B. Mass (Nucleon) Number

- The mass number is the total number of protons and neutrons in an atom's nucleus.
- It is not found on the periodic table; it's often written as a superscript.
- Formula: Mass Number = Protons + Neutrons
- Example: Carbon-12 has 6 protons and 6 neutrons, so its mass number is 12.

C. Atomic Mass

- The average mass of all the naturally occurring isotopes of an element, weighted by their abundance.
- It is usually shown on the periodic table.
- It is not always a whole number.
- Example: The atomic mass of chlorine is about 35.5 due to a mix of Cl-35 and Cl-37 isotopes.

D. Atomic Mass Unit

- An atomic mass unit (amu) is a unit of mass used to express atomic and molecular weights.
- 1 amu is defined as one-twelfth the mass of a carbon-12 atom.
- Also written as u or Da (Dalton).
- 1 amu ≈ 1.66×10^{-27} kg

3.1.3 Calculate The Atomic Number, Mass Number, Number Of Electrons And Neutrons Of Atoms And Ions

Number of Neutrons (N) = mass number (A) – atomic number (Z)

Number of electrons = atomic number

3.1.4 Draw The Atomic Structure Of The First Twenty Elements Of The Periodic Table And Their Ions (Cations And Anions) Using Their Mass Number And Atomic Number

Protons and neutrons will be shown in the center of atom. Electrons will revolve around the nucleus in shells:

- K-Shell – maximum 2 electrons allowed
- L-Shell – maximum 8 electrons allowed
- M-Shell – maximum 18 electrons allowed
- N-Shell – maximum 32 electrons allowed

This configuration follows a general rule of $2n^2$ where n is number of shells.

3.2 Isotopes

3.2.1 Define The Following Terms:

A. Isotopes

- Isotopes are atoms of the same element (same number of protons) but with different numbers of neutrons, and therefore different mass numbers.
- Example: Carbon-12 and Carbon-14 are isotopes of carbon. Both have 6 protons, but Carbon-12 has 6 neutrons, and Carbon-14 has 8.

B. Average Atomic Mass

It is the weighted average of all naturally occurring isotopes of an element. It is sometimes used in mathematical calculations or charts.

C. Relative Atomic Mass Based On C-12 Scale

It is the average mass of an atom of an element compared to 1/12 the mass of a carbon-12 atom. It takes into account the different isotopes of an element and their natural abundances. It has No units — it's a relative value.

$$Relative\ Atomic\ Mass = \frac{(m1 \times a1) + (m2 \times a2) + \cdots}{100}$$

Where:

- m1, m2, ... = mass numbers of isotopes
- a1, a2, ... = percentage abundances of isotopes

D. Radioactive Isotopes

Radioactive isotopes are unstable forms of elements that have the same number of protons but a different number of neutrons, and they emit radiation as they break down into more stable atoms.

Carbon-14 (^{14}C) — used in carbon dating

Iodine-131 (^{131}I) — used in medical diagnostics

E. Radioactivity

Radioactivity is the spontaneous emission of radiation (alpha, beta, or gamma rays) from the nucleus of an unstable atom. This process continues until the atom becomes stable.

3.2.2 Determine The Number Of Protons, Neutrons And Electrons In Different Isotopes

Isotope	Protons	Neutrons	Electrons	Notes
^{1}H	1	0	1	Protium (most common H)
^{2}H	1	1	1	Deuterium (heavy hydrogen)
^{3}H	1	2	1	Tritium (radioactive)
^{12}C	6	6	6	Most common carbon isotope
^{13}C	6	7	6	Stable isotope
^{14}C	6	8	6	Radioactive (used in dating)
^{16}O	8	8	8	Most abundant oxygen isotope
^{17}O	8	9	8	Rare stable isotope
^{18}O	8	10	8	Stable isotope
^{35}Cl	17	18	17	About 75% of chlorine
^{37}Cl	17	20	17	About 25% of chlorine
^{238}U	92	146	92	Most abundant uranium
^{235}U	92	143	92	Used in nuclear reactors

3.2.3 Calculate The Relative Atomic Masses Of Chlorine And Boron By Using The Mass Number And Natural Abundance Of Their Isotopes

Chlorine (Cl)

Chlorine has two common isotopes:

- Cl-35 (mass = 35, abundance = 75%)
- Cl-37 (mass = 37, abundance = 25%)

$$Relative\ Atomic\ Mass = \frac{(35 \times 75) + (37 \times 25)}{100} = 35.5$$

Boron (B)

Boron has two common isotopes:

- B-10 (mass = 10, abundance = 20%)
- B-11 (mass = 11, abundance = 80%)

$$Relative\ Atomic\ Mass = \frac{(10 \times 20) + (11 \times 80)}{100} = 10.8$$

3.2.4 Explain The Role Of Isotopes

Radioactive isotopes are useful in medical imaging. Doctors use them to diagnose the disease by injecting the patient with a small amount of radioactive fluid. Technetium- 99m is used for diagnostic imaging across human organs like brain, lungs, etc. Doctors use a special camera to watch how the radioactive fluid moves.

Radiocarbon dating is a method for finding out the age of an historical object containing organic material with the help of radioactive isotope of carbon ^{14}C. The method involves measuring the proportion of ^{14}C in a sample from a dead plant or animal like a piece of wood or a bone which provides information that can be used to calculate when an animal or plant died. The older the sample is, the less ^{14}C is to be detected.

Radioactive isotopes are used to test the strength of metals and concrete mixture. They are used to generate cheap nuclear power and to find oil fields. In medicine they are used to diagnose and treat many medical conditions and diseases, including cancer and thyroid disorders.

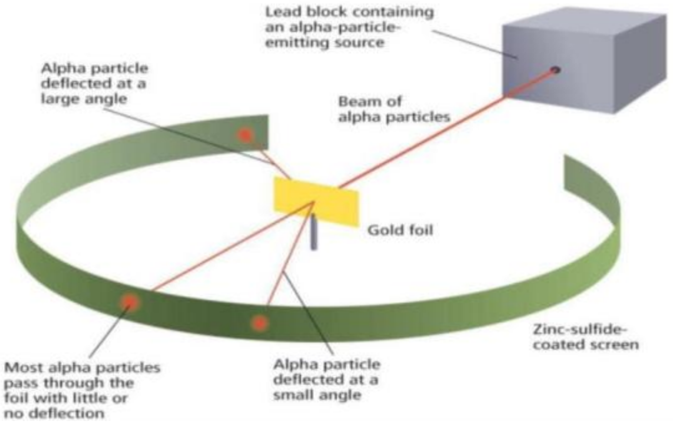
Isotope	Use in Medical Imaging
Iodine-131 (^{131}I)	Used to diagnose and treat thyroid conditions (thyroid absorbs iodine)
Sodium-24 (^{24}Na)	Traces blood flow and circulation problems
Technetium-99m	Most widely used — images of bones, heart, lungs, kidneys
Thallium-201 (^{201}Tl)	Used in stress tests to image blood flow in the heart
Arsenic-74 (^{74}As)	Helps locate tumors and track organ function
Cobalt-60 (^{60}Co)	Used in radiotherapy to kill cancer cells
Xenon-133 (^{133}Xe)	Used in lung scans to image air flow

3.3 Models to Understand the Structure of an Atom

3.3.1 Explain Rutherford's Experiment Leading To Discovery Of The Atomic Nucleus

Lord Rutherford in 1911, carried out series of experiments and proposed a new model for the atom.

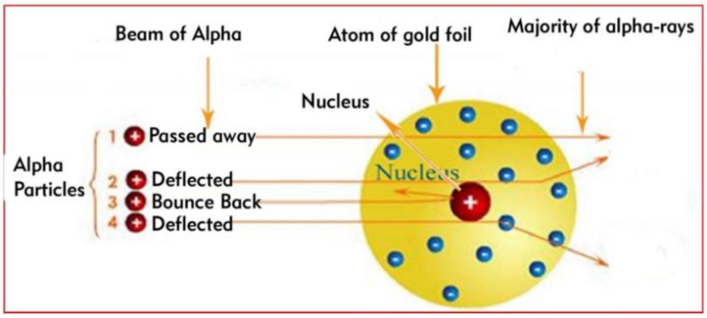
He took a thin sheet of gold foil and bombarded it with alpha (α) particles obtained from a radioactive element (Like Polonium). These rays were scattered after passing through the foil and examined on a zinc sulphide (ZnS) screen.



Observations

1. Most of the particles passed straight and undeflected through the sheet and produced illumination on the zinc sulphide screen.
2. Very few alpha (α) particles undergo small and strong deflection after passing through gold sheet.
3. A very few alpha (α) particles (one out of 8000) retraced their path.

Conclusion



1. According to Rutherford an atom consists of two parts nucleus and extra nuclear part.
2. Majority of the alpha particles passed straight line and un-deflected, shows that most volume occupied by atom is empty.
3. Alpha particles are positively charged and their deflection indicates that the center of atom has a positive charge, which is named as nucleus.
4. The mass is concentrated in the nucleus and the electrons are distributed outside the positively charged nucleus.
5. The electrons are revolving around the nucleus in extra nuclear part called orbits.

Rutherford Postulates

1. An atom consists of positively charged, dense and very small nucleus containing protons and neutrons. The entire mass is concentrated in the nucleus of an atom.
2. The nucleus is surrounded by large empty space which is called extra nuclear part where probability of finding electron is maximum.
3. The electrons are revolving around the nucleus in circular paths with high speed (Velocity). These circular paths are known as orbits (Shells).
4. An atom is electrically neutral because it has equal number of protons and electrons.
5. The size of the nucleus is very small as compared to the size of its original atom.

3.3.2 Explain The Defects Of Rutherford's Atomic Model

1. Rutherford did not explain the stability of an atom.
2. In Rutherford atomic model the negatively charged electrons revolve around the nucleus in circular path and emits energy continuously. Due to continuous loss of energy ultimately it must fall into the nucleus.
3. If the revolving electron emits energy, then there would be a continuous spectrum but in contrast to it we get line spectrum from the atoms of elements.

3.3.3 Describe The Main Points (Postulates) Of Bohr's Atomic Model

1. The atom has fixed orbits in which negatively charged electrons are revolving around the positively charged nucleus.
2. These orbits possess certain amount of energy which are called shells and named as K, L,M, N shells.
3. The energy levels are represented by an integer (n= 1, 2, 3....) known as quantum number, this quantum range starts from nucleus side, where n= 1 is lowest energy level.
4. Electrons are revolving in particular orbits (ground state) continuously, but they do not emit energy.
5. When electron absorbs energy, it jumps from lower energy level (E_1) to higher energy level (E_2).
6. When electrons jump from higher energy level (E2) to lower energy level (E 1), it emits energy. The electron present in the orbit which is closest to the nucleus, has minimum energy and it is called the ground state of the atom. The orbits that are further away from the nucleus possess successively greater energy.
7. The electron is not allowed to occupy a space in between the orbits.
8. Since electron present in each shell has a fixed energy, these shells are also named as energy levels.
9. Each shell is further sub-divided into sub-shells or orbitals. The number of sub-shells present in a shell is equal to the value of 'n' for that shell.

3.4.1 Differentiate Between Shells And Sub-Shells Of An Atom

Feature	Shells (Energy Levels)	Sub-shells (Sub-energy Levels)
Definition	Main energy levels where electrons are found	Divisions within shells where specific types of electrons exist
Notation	Represented by numbers (n = 1, 2, 3...) or letters (K, L, M...)	Represented by letters: s, p, d, f
Number per atom	Limited by principal quantum number n	Number of sub-shells = n (same as shell number)
Electron capacity	Shell capacity = $2n^2$ electrons	Each sub-shell has a specific capacity: s = 2, p = 6, d = 10, f = 14

Example: For the element Oxygen (Atomic number = 8):

- Shells: 2 electrons in the K-shell (n=1), 6 in the L shell (n=2)
- Sub-shells in K-shell: 2 electrons in 1s
- Sub-shells in L-shell: 2 electrons in 2s, 4 electrons in 2p

3.5 Electronic Configuration

3.5.1 Deduce The Electronic Arrangement (In Shells) And Electronic Configuration (Of Sub-Shells) Of The First Twenty Elements And Their Ions

Element	Atomic Number	Shell Arrangement	Sub-shell Configuration	Common Ion	Ion Configuration
H	1	1	$1s^1$	H^+	— (no electrons)
He	2	2	$1s^2$	—	—
Li	3	2,1	$1s^2\ 2s^1$	Li^+	$1s^2$
Be	4	2,2	$1s^2\ 2s^2$	Be^{2+}	$1s^2$
B	5	2,3	$1s^2\ 2s^2\ 2p^1$	—	—
C	6	2,4	$1s^2\ 2s^2\ 2p^2$	—	—
N	7	2,5	$1s^2\ 2s^2\ 2p^3$	N^{3-}	$1s^2\ 2s^2\ 2p^6$
O	8	2,6	$1s^2\ 2s^2\ 2p^4$	O^{2-}	$1s^2\ 2s^2\ 2p^6$
F	9	2,7	$1s^2\ 2s^2\ 2p^5$	F^-	$1s^2\ 2s^2\ 2p^6$
Ne	10	2,8	$1s^2\ 2s^2\ 2p^6$	—	—
Na	11	2,8,1	$1s^2\ 2s^2\ 2p^6\ 3s^1$	Na^+	$1s^2\ 2s^2\ 2p^6$
Mg	12	2,8,2	$1s^2\ 2s^2\ 2p^6\ 3s^2$	Mg^{2+}	$1s^2\ 2s^2\ 2p^6$
Al	13	2,8,3	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$	Al^{3+}	$1s^2\ 2s^2\ 2p^6$
Si	14	2,8,4	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$	—	—
P	15	2,8,5	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$	—	—
S	16	2,8,6	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$	S^{2-}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Cl	17	2,8,7	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$	Cl^-	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Ar	18	2,8,8	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	—	—
K	19	2,8,8,1	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	K^+	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Ca	20	2,8,8,2	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$	Ca^{2+}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$