Engineering of Radiation Instrumentation

Atomic Structure and Atomic Radiation :

1. Basic Atom

The atom :
Is the smallest unit of matter that retains the chemical properties of an element.

Atom consists of a **nucleus** and **orbital electrons**.

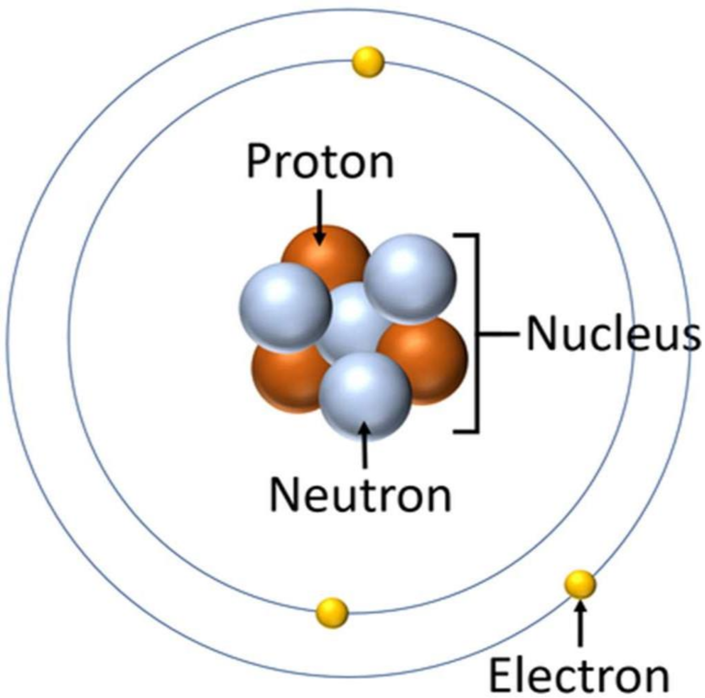


Figure (1): The atomic structure

The nucleus : Is a heavy central mass with a positive charge.

The nucleus accounts for most of the mass of an atom, but is small in diameter compared to the overall size of the atom.

Outside the nucleus is a cloud of **electrons** - small, negatively charged particles that are attracted to the atom by the opposing charge of the nucleus (coulomb forces).

2. Protons and Neutrons (Nucleons)

Nucleons : Particles within the nucleus refer to protons and neutrons .

- The **nucleus** is composed of **protons** and **neutrons**.
  - **Protons** have a positive (+ 1) charge with a mass of about 1.007 atomic mass units which is equal to (1.673 x 10^{-24} grams).

- **Neutrons** are neutral, or uncharged. Neutrons have a mass slightly larger than the proton, about 1.008 atomic mass units (1.675 x 10^{-24} grams). Although the coulomb forces between protons in the nucleus tend to push the nucleus apart, the nucleus is held together by the strong nuclear force that operates at extremely short distances. The strong nuclear force is believed to result from attractions between even smaller particles (quarks) that compose the neutrons and protons.

3. Electrons:

Electrons : Are very small, negatively charged particles .

- **The charge of the electron** (- 1) **is equal in magnitude to the charge of the proton** but is negative instead of positive.
- **In an uncharged atom**,

No. of orbital electrons (e) = No. of protons within the nucleus (p) or **(p = e)**

- When the number of orbital electrons does not equal the number of protons in the nucleus, an overall imbalance of charge exists for the atom.
- **A charged atom is known as an ion**. Ions readily form chemical bonds with other ions of opposing charge.

- Although electrons exist in a cloud around the nucleus, it is useful to describe this arrangement as a series of **energy levels, called shells**. Within **each shell are subgroups of electrons, called orbits.** An atom may have a number of possible energy states, which correspond to different arrangements of electrons in the shells.

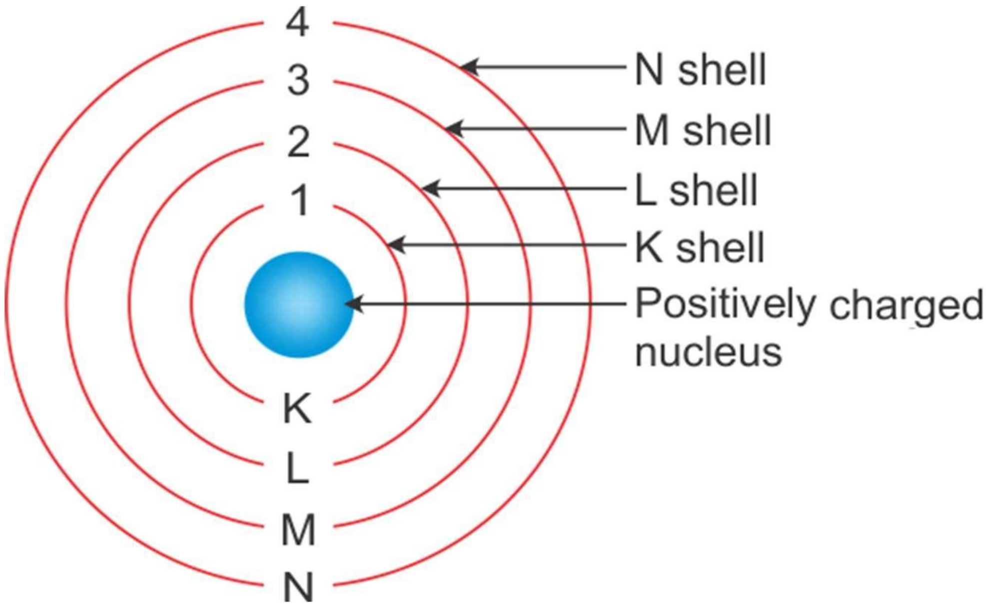


Figure (2):The energy levels (shells) of the atom

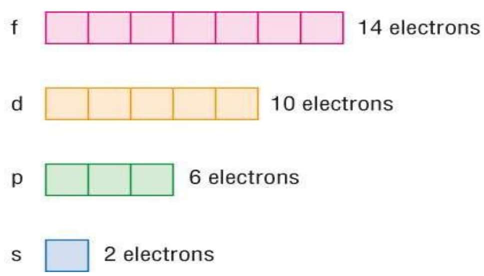


Figure (3): The s,p,d and f orbitals (subshells)

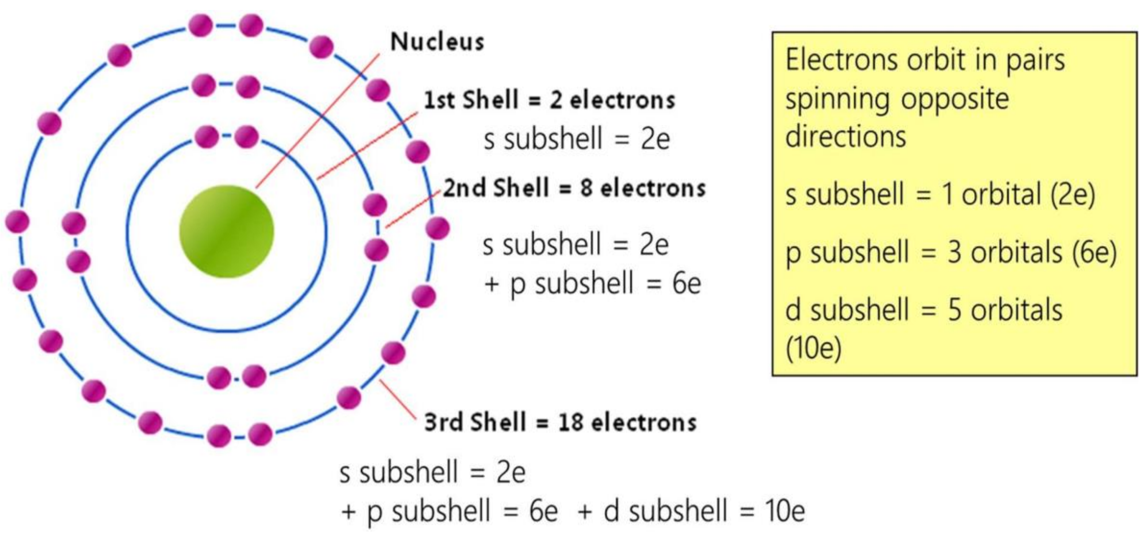


Figure (4):The shells and subshells in atom

4. Atomic Number, and Atomic Mass Number

- The **number of protons** in the nucleus of an atom is known as the **atomic number** and is represented by the symbol **Z**.

- Atoms are grouped into **elements** according to their atomic number.
- Each element is also represented by a chemical name.

- The **total number of protons and neutrons** in the nucleus of an atom is called the **atomic mass number (A)** (or, **mass number**).

- The atomic mass number should not be confused with the atomic mass.

- **Atomic mass** describes the **relative mass of the atom** (including orbital electrons).

- The scale for atomic mass is fixed so that it equals 12.000 amu for an atom having 6 protons, 6 neutrons and 6 electrons. This atom (Carbon-12) also has an atomic mass number of 12. For all other atoms, the atomic mass and atomic mass numbers differ slightly in magnitude.
- The **number of neutrons** in an atom is called the **neutron number** and is represented by the symbol **N**.

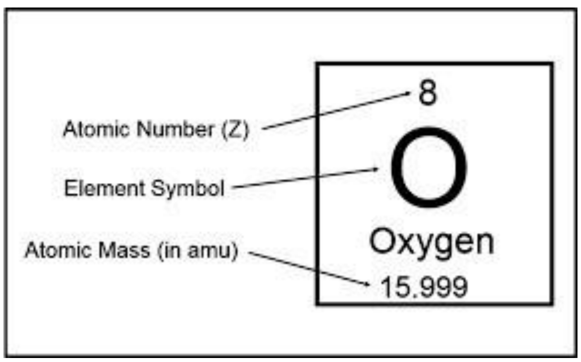


Table 1 illustrates the relation of A, Z and N for several atoms.

Table 1 – Examples of Z, N, and A

	Atomic No. Z (Protons)	*Neutron No. N (Neutrons)*	*Mass No. A (Nucleons)*
Hydrogen-1	**1**	**0**	**1**
Carbon-12	**6**	**6**	**12**
Iron-56	**26**	**30**	**56**

Bohr’s Model of the Hydrogen Atom

- ❖ **Niels Bohr** in 1913 when **he presented a new model** of the hydrogen atom that circumvented the difficulties of Rutherford’s planetary model.
- ❖ **Bohr** applied **Planck’s ideas of quantized energy levels** to **Rutherford’s orbiting atomic electrons.**
- ❖ **Bohr** combined ideas from **Planck’s original quantum theory**, **Einstein’s concept of the photon**, **Rutherford’s planetary model of the atom**, and **Newtonian mechanics** to arrive at a **semiclassical model** based on some revolutionary ideas.
- ❖ The postulates of the Bohr theory as it applies to the hydrogen atom are as follows:

1. The electron moves in **circular orbits** around the proton under **the influence of the electric force of attraction** as shown in Figure 5.

2. Only certain electron orbits are stable. When in one of these stationary states, as Bohr called them, **the electron does not emit energy in the form of radiation**, even though it is accelerating. Hence, the total energy of the atom remains constant and classical mechanics can be used to describe the electron’s motion. Bohr’s model claims that the centripetally accelerated electron does not continuously emit radiation, losing energy and eventually spiraling into the nucleus, as predicted by classical physics in the form of Rutherford’s planetary model.

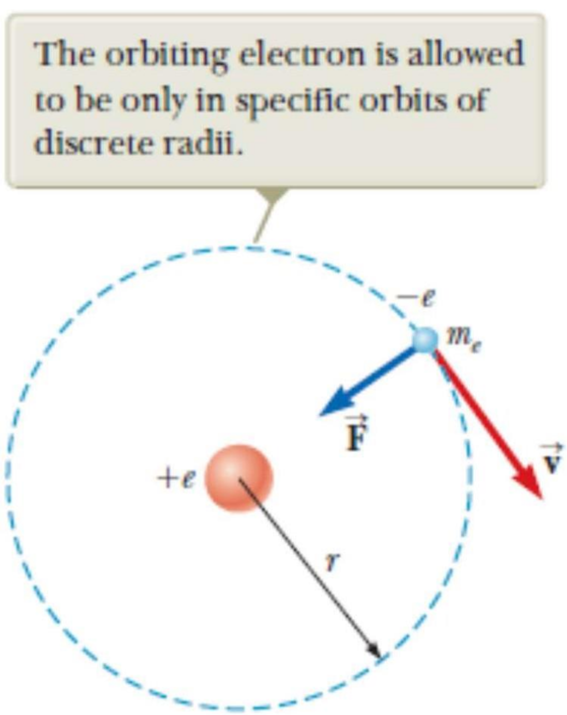


Figure (5):diagram representing Bohr’s model of the hydrogen atom

3. **The atom emits radiation** when **the electron makes a transition from a more energetic initial stationary state to a lower-energy stationary state**.

This transition cannot be visualized or treated classically. In particular, the frequency **υ** of the photon emitted in the transition is related to the change in the atom's energy and is not equal to the frequency of the electron's orbital motion. The frequency of the emitted radiation is found from the energy-conservation expression

$$\mathbf{E_i - E_f = h\upsilon} \qquad \textbf{(eq. 1)}$$

Where: E_i is the energy of the initial state (high energy level),

E_f is the energy of the final state (low energy level), $E_i > E_f$.

υ is the frequency the emitted photon.

h is planck constant $6.626*10^{-34}$ j.s

4. The size of an allowed electron orbit is determined by a condition imposed on the electron's orbital angular momentum: the allowed orbits are those for which **the electron's orbital angular momentum about the nucleus** is **quantized and equal to an integral multiple of ђ=h/2π**

$$\mathbf{m_e v r = n\hbar} \qquad \textbf{(eq. 2)}$$

where $\mathbf{m_e}$ is the electron mass , **v** is the electron speed in its orbit and **r** is the orbital radius .

n=1,2,3,…… for first ,second and third orbitals

$$r_n = n^2 a_0 = n^2(0.0529\ \text{nm}) \quad n = 1, 2, 3, \ldots$$

(eq. 3) Radii of Bohr orbits in H_2

What is Radiation?

Radiation is energy that comes from a source and travels through space at the speed of light.

- This energy has an **electric field** and a **magnetic field** associated with it, and has wave-like properties. You could also call radiation "**electromagnetic waves**".

- The **radiation** has a particular **wavelength** and **frequency** that defines it energy.

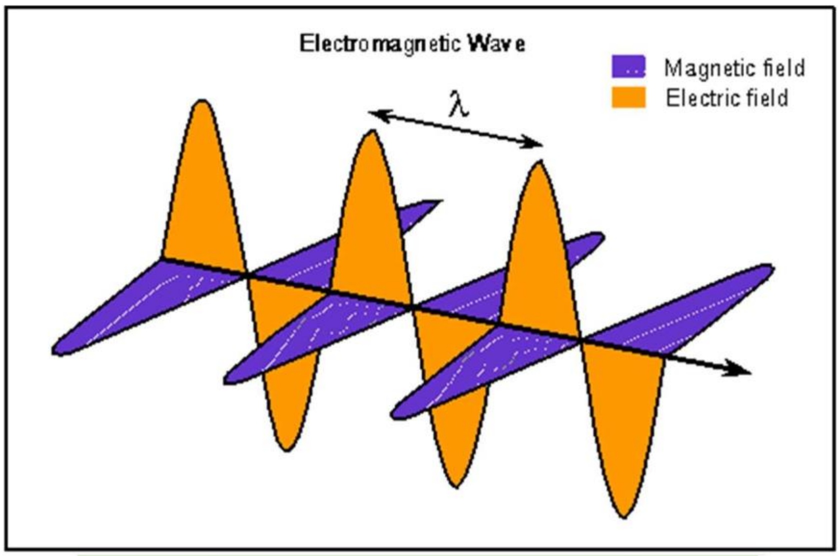


Figure (6): Electromagnetic wave of the radiation

Photon Energy : -

(Planck Equation)

$$E = h\nu = \frac{hc}{\lambda}$$

Where,

E = Photon Energy

h = Planck constant $= 6.6261 \times 10^{-34}$ J*S

c = Speed of light $= 3 \times 10^{8}$ m/s

λ = Photon wavelength

v = Photon frequency

Note:
The energy of the emitted photon = the energy of electromagnetic radiation

❖ The **electromagnetic spectrum** is the range of **frequencies (the spectrum) of electromagnetic radiation** and their respective wavelengths and photon energies.

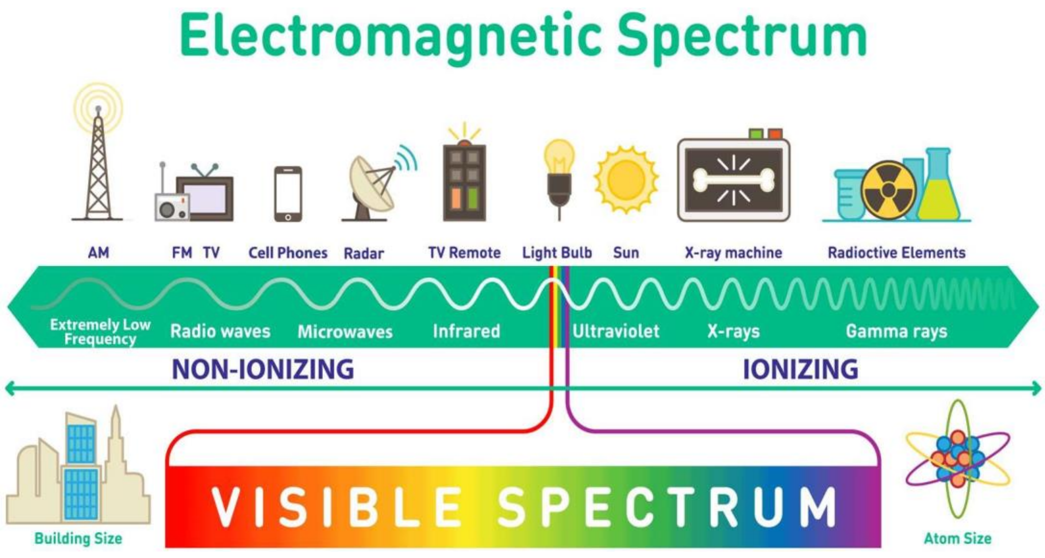


Figure (7): The Electromagnetic Spectrum

❖ The most common form of radiation is the **visible light**, we can detect this radiation with our eyes. The only difference between various colors of light, red, yellow, green, blue, and purple is in their wavelength or frequency, or in other words in their energy. **Red light**, for example, has **less energy** than **purple light**.

❖ There is a wide range of electromagnetic radiation in nature. The visible part of the spectrum is only a tiny part of this wide range of energies.

❖ As **we move down in frequency from red light**, there are other familiar forms of electromagnetic radiation:

- Infrared
- Microwaves
- Signals from our cell phones
- Radio waves

These are all forms of radiation that are **invisible to our eyes** and that have **less energy than visible light**.

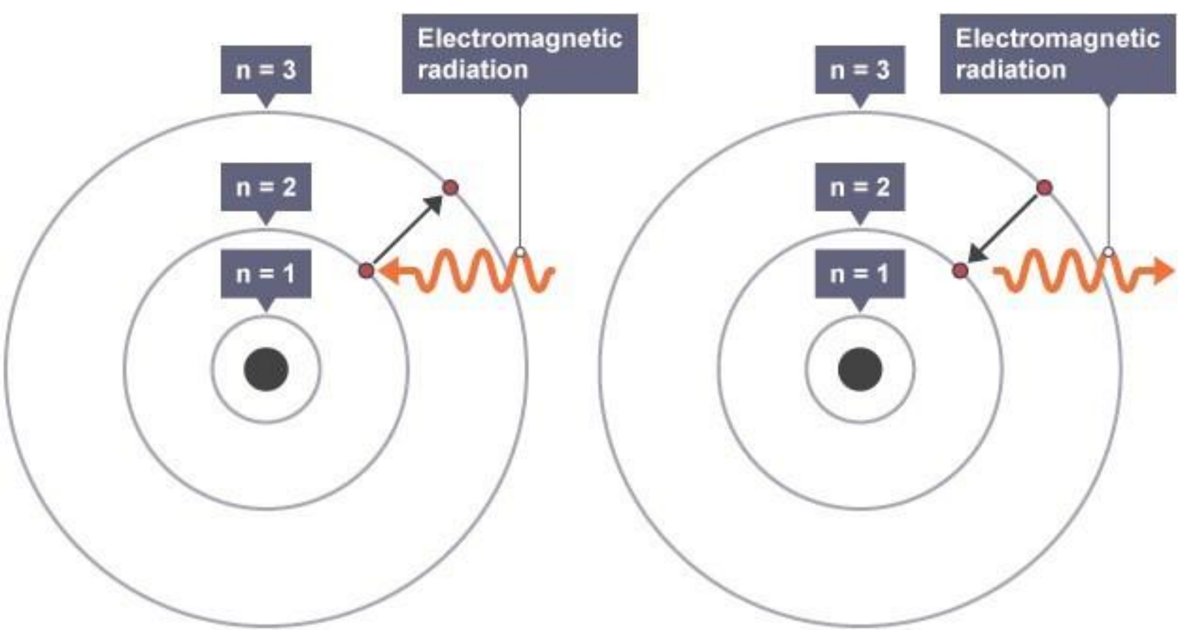
❖ As we move up in frequency from purple light, there are

- Ultraviolet (UV) radiation
- X-rays
- Gamma rays

These are all forms of radiation with **energies much higher than visible light**.

How does the atomic radiation happen?

- **An atom may have a number of possible energy levels**, which correspond to different arrangements of electrons in the shells. **Electrons normally occupy the lowest energy levels in the atom**, with successive shells and orbits filled in a complex manner.
- **If an atom absorbs energy**, **through heating**, **from electricity**, or **by absorbing electromagnetic radiation**, **the electrons at a particular level can be pushed up to higher levels (at bigger distances from the nucleus)**. In time, they jump back down to a lower level radiating energy in the form **of electromagnetic waves (emission).**



Figure(8):The absorption and emission in atom.

The frequency of the radiation depends on **the size of the jump that the electron makes**. **Sodium** has energy levels that cause the electrons to emit **yellow light** as they drop from one to the other. **The energy levels in mercury** are further apart and electrons emit **ultraviolet radiation** when they jump back down.

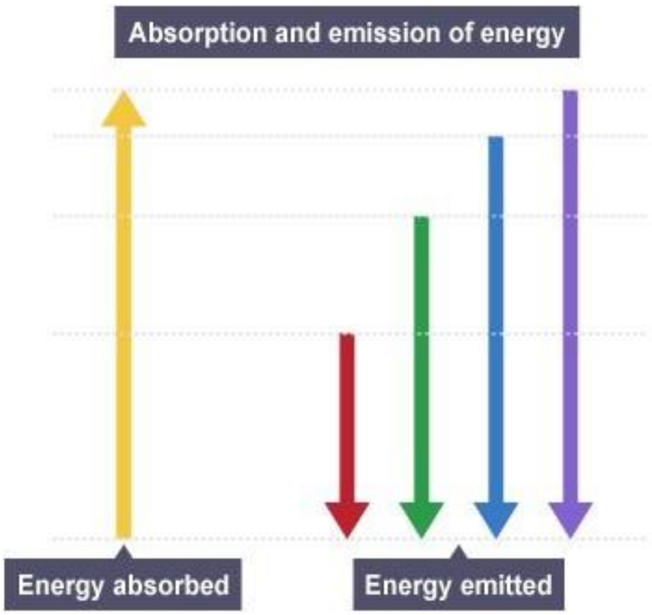


Figure (9): In the diagram, the colors of the 'down arrows' represent the color of the light emitted as the electron makes that drop.

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