SENIOR FIVE TERM 1

TOPIC 2/3: Atomic and Electronic Structure

Competency: The learner deduces electronic configurations, evaluates their implications for chemical properties and bonding, and synthesises models to predict atomic behaviour in various contexts

Atomic structure

This chapter teaches the properties of the components of an atom. The placement of electrons is specifically emphasized because this determines the physical and chemical properties of an element.

Matter

It is anything that occupies space; it is composed of discrete particles called atoms.

An atom is the smallest indivisible electrically neutral particle of an element that can take part in a chemical reaction.

The structure of the atom

The atom is composed of three basic subatomic particles, namely the electron, the proton and the neutron; the characteristics of which are given in a table 1 below:

Table 1: The three main subatomic particles

Particle	Symbols	Charge	Mass
Electron	e	-1	1/1837
Proton	p	+1	1
Neutron	n	No charge	1

All atoms of the same element contain the same number of protons (usually equal to the number of electrons). The number of protons in an atom is characteristic of an element and is called the **atomic number**. **The atomic mass** is the sum of the number of protons and neutrons in an atom (element).

Atoms with the same atomic number but different atomic mass (due to the difference in the number of neutrons in an atom) are called **isotopes.**

The arrangement of electrons in atoms

An atom is composed of a nucleus (proton and neutrons) situated at the center of the atom and surrounded by a system of electrons. The electrons rotate around the nucleus in definite orbital called energy levels or **quantum shells**; which are specified by giving them numbers, i.e., 1, 2, 3, 4, --------- or letters (capital) K, L, M, ------------.

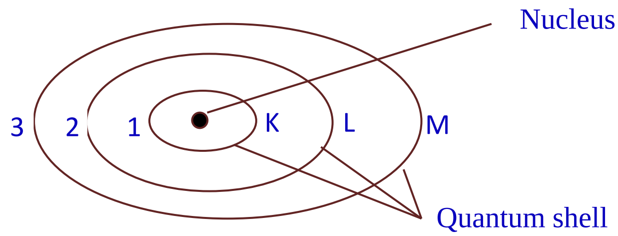


Fig. 1: A drawing showing quantum shells in an atom

The numbers, 1, 2, 3 ----------- are called **principal quantum number**. Each energy level is associated with a definite amount of energy and the energy of these levels increases in the order K<L<M<N. An electron in any of these energy levels is associated with a definite quantity of energy.

If an extra energy is supplied to the electron in ground state (this is usually done by heating, by collision with a fast-moving particle or by electrical discharge), the electron absorbs this extra energy and jumps to one of the higher energy levels, Fig. 2. In this situation, the electron is said to be **excited**.

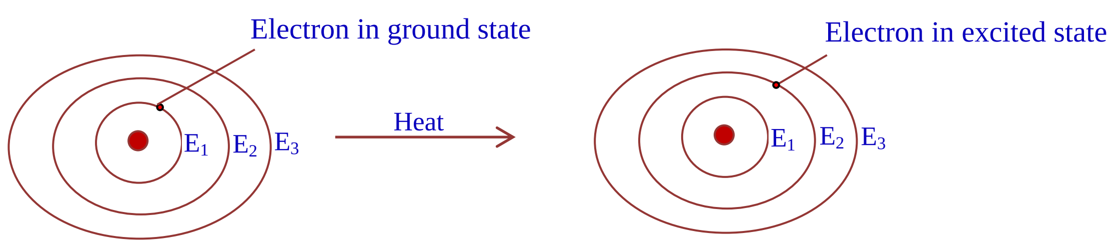


Fig. 2. Excitation of an electron

If the electron falls back to the lower energy level, it emits energy in form of radiation. The energy absorbed or emitted is equal to the energy difference between the two levels.

Atomic emission spectra

Atomic spectrum is a characteristic pattern of radiations absorbed or emitted by a substance and it is one of the physical features that provides evidence for existence of energy levels within an atom.

(a) Absorption spectra

All atoms and molecules absorb light of certain wavelength. When light is passed through a gaseous substance, at low pressure black lines appear in the spectrum, where these wavelengths of light have been absorbed by substance. The pattern of frequencies of light absorbed by a substance is called its **absorption spectrum**. The absorption spectrum is characteristic and can be used to identify a substance.

(b) Atomic emission spectrum or line spectrum

If atoms and molecules are heated to sufficiently high temperature, they emit light of certain wave lengths, for example, sodium gives out a yellow flame, Fig. 3.

.

Fig. 3

The pattern of frequencies of light emitted by a substance is called its **atomic emission spectrum** or **line spectrum**. Each line or color corresponds to a definite wavelength of radiation. If the frequency of the radiation is in the visible region of the electromagnetic spectrum; then it is seen as a colored light on a black background.

Applications of atomic emission spectra

(i) **Chemical analysis**: Atomic emission spectroscopy (AES) is used to identify and quantify elements in samples. For example, detecting trace metals in alloys or solutions.

(ii) **Advertising**: emission spectra are used for street light advertising due their attractive colors as shown in the Fig. 4

Fig. 4

(iii) **Astronomy**: Spectral lines from stars and galaxies reveal their chemical composition, temperature, and motion. Hydrogen and helium lines are key in studying stellar evolution.

(iv) **Environmental monitoring**: Used to measure pollutants such as heavy metals in water, soil, and air. AES helps track contamination and ensure safety standards.

(v) **Medical diagnostics**: Plasma emission techniques detect trace elements in biological samples (like blood or tissue), aiding in disease diagnosis and nutritional studies.

(vi) **Industrial quality control**: Emission spectra ensure correct composition in metals, semiconductors, and ceramics. Industries use it to maintain product standards.

(vii) **Forensic science**: Identifies unknown substances or contaminants in crime scene samples by analyzing their emission spectra.

(viii) **Education and research**: Demonstrates quantum theory concepts, electron transitions, and energy levels in physics and chemistry labs.

- **Advantages:** Highly sensitive, can detect trace elements, provides rapid results.
- **Challenges:** Requires calibration, expensive instrumentation, and sometimes destructive sampling.

The hydrogen atomic spectra

(i) Absorption spectrum

It is observed as black lines on a black background when white light is shone through gaseous hydrogen at low pressure. The absorption spectrum is caused by hydrogen atoms absorbing energy corresponding to certain wavelengths from light.

(ii) The hydrogen emission spectrum

The **hydrogen emission spectrum** is a set of discrete lines observed as a pink glow produced when excited hydrogen atoms release energy as electrons fall back to lower energy levels.

Viewed through a spectrometer, the emission is seen to be a number of separate sets of lines or series of lines. The series of line are named after their discoverers as shown in Fig 5.

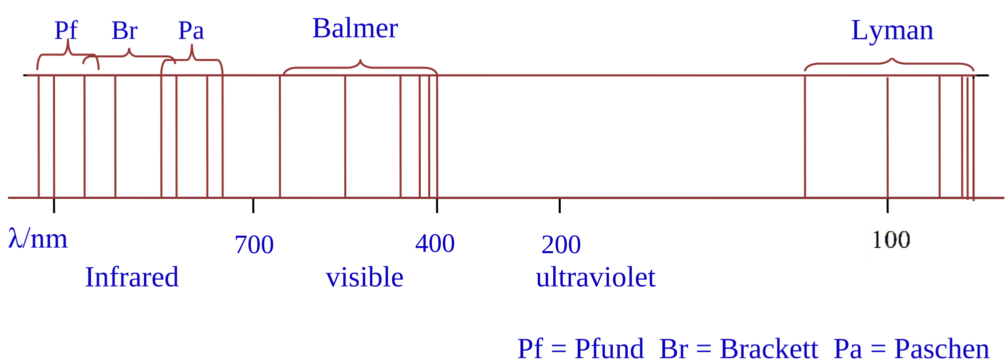


Fig. 5 Emission spectrum of hydrogen

In each series, the interval between the frequencies of the lines become smaller and smaller towards the high frequency end of the spectrum until the lines run together or converge to form a continuum of light. For example, Balmer series, in the visible part of the spectrum is shown in Fig. 6

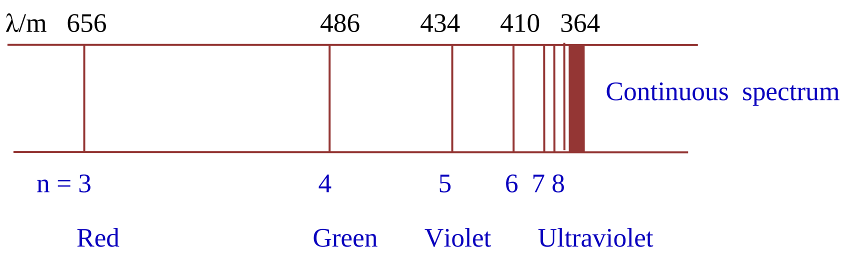


Fig. 6. Balmer series of Hydrogen

Cause of hydrogen spectrum

Electrons in hydrogen atoms absorb energy and move to higher energy levels. When they return to lower levels, they emit photons of specific energies or definite frequencies of definite characteristic colors. Energy absorbed is equal to the energy difference between two energy levels.

Relationship between energy and frequency wavelength absorbed or emitted.

The discrete amount of energy (quantum) absorbed or emitted by an atom is proportion to the frequency of radiation as follows.

E = hf (Plank's equation)

where E is the energy in joules, f, is the frequency of radiation (velocity of light/wavelength) and h is called Plank's constant. Its value is 6.625 x 10^{-34}Js

When the lines corresponding to a particular series are examined, they are seen that they all fit in the equation.

$$\frac{1}{\lambda} = RH\left(\frac{1}{n^2} - \frac{1}{m^2}\right)$$

where λ is the wavelength of a particular line, RH is the Rydberg's constant and, n and m are integers, (m > n).

The complete atomic spectrum of hydrogen is resolved into five definite series characterized by one value of n for each series and with varying values of m.

Table 2: Tabular representation of atomic hydrogen spectrum.

Layman series	n = 1	m = 2, 3, 4, etc.	in the ultra violet region.
Balmer series	n = 2	m = 3, 4, 5, etc.	in the visible region
Paschen series	n = 3	m = 4, 5, 6, etc.	in the infrared region.
Brackett series	n = 4	m = 5, 6, 7, etc.	in the far infrared region.
Pfund series	n = 5	m = 6, 7, 8, etc.	in the far infrared region.

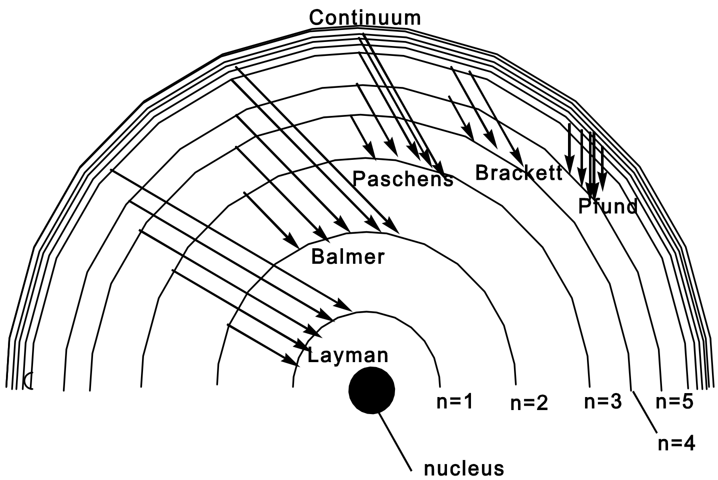


Fig. 7. The electron transition between energy levels or orbitals in each series.

Ionization

Ionization occurs when an electron jump beyond the highest energy level and does not return to an atom. The energy required to excite an electron from its ground state above the highest energy level is the **ionization energy** for that electron.

How does hydrogen spectrum provide evidence for existence of energy levels in an atom

Being a line spectrum indicates that only definite energy absorption and emission are permissible during electron transition within an atom; proving existence of energy levels within an atom.

Examples 1

(a) Calculate the wave number for the 1st and 3rd line in the layman's series

(RH = 109678 cm^{-1})

Solution:

But for the Layman's series n=1 and first line m=2

$$Wave\ number, \frac{1}{\lambda} = RH\left(\frac{1}{n^2} - \frac{1}{m^2}\right)$$

$$Wave\ number, \frac{1}{\lambda} = 109678\left(\frac{1}{1^2} - \frac{1}{2^2}\right) = 82259 cm^{-1}$$

or $\lambda = 1.216 \times 10^{-5}$cm

For the third line in Layman's series m = 4

$$Wave\ number, \frac{1}{\lambda} = 109678\left(\frac{1}{1^2} - \frac{1}{4^2}\right) = 102823 cm^{-1}$$

or $\lambda = 9.725 \times 10^{-6}$cm

(b) Calculate the wave number for the line with n=1 and m=∞ (RH = 109678 cm^{-1}

$$Wave\ number, \frac{1}{\lambda} = 109678\left(\frac{1}{1^2} - \frac{1}{\infty^2}\right) = 109678 cm^{-1}$$

Thus, the wave number 109678 corresponds to the highest emission line in the Layman's series

Trial 2.1

2. (a) Describe the spectrum of a hydrogen atom
Use a diagram to illustrate your answer (7marks)

(b) Explain how the spectrum of a hydrogen atom:

(i) is formed (4marks)

(ii) provides evidence for existence of energy levels in atoms (7marks)

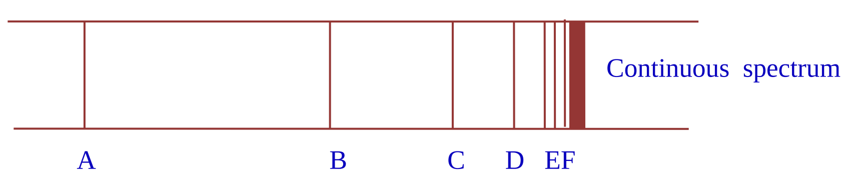
(c) The frequency of hydrogen at the point of ionization is 32.8×10^{14} Hz

Calculate the ionization energy of hydrogen

(Plank's constant = 6.6×10^{-34}Js) (2marks)

3. Hydrogen spectrum consists of several series of lines. Each line represents electron transition between energy levels characterized by different value of n.

Part of the highest energy Layman's series in the hydrogen atomic spectrum is shown in figure below. A is the first line of this series



(a)(i) On the diagram draw an arrow to show the direction of increasing energy and label it the arrow energy.

arrow from left to right

(ii) Draw another arrow of increasing frequency and label it frequency

arrow from left to right

(b) Why does the spectrum consists of lines

Because electron transitions involve definite amount of energy.

(c) Name the symbol n

Principal quantum number

(d) What do the transition in the same series all have in common

All involve transitions from higher levels down to the same energy level

(e) When, if any, of the lines A to F shown above correspond to each of the following transition? If none of the lines correspond, answer 'none. In each case explain your answer

(i) transition N=3 to n=1

C

(ii) the transition n=3 to n=2

None because in Lyman's series n = 1

(iii) The transition n=1 to n=3

None because in the spectrum electrons drop from high to low energy levels

Subdivision of the main energy levels

The main energy levels of an atom designated by the principal quantum number, n = 1, 2, 3, 4… or by letter K, L, M, N…, are further subdivided into sub- energy levels; for instance, the 8 electrons in the L and M shells of potassium (n = 2 and 3 respectively), are distributed between two sub-levels containing 2 and 6 electrons respectively. The four major set of energy levels that characterize an electron are described below:

1. The **principal quantum number, n,** has integral values 1, 2, 3, 4, etc. These numbers define the energy level
2. 16 of an electron in the atom, e.g., principal quantum number 1; represent the orbital nearest to the nucleus and this orbital has the lowest energy. An electron with the largest value of n, has the most energy and furthest from the nucleus. It is the one that requires the least input of energy to ionize, i.e., it is one most readily removed from an atom.

2. The **azimuthal (subsidiary) quantum number (l**), has integral values from 0, 1, 2, --- (n-**l**). This quantum number represents the associated orbitals in a given principal quantum shell. This quantum number is however, usually denoted by using symbols, s, p, d, f, ... instead of numbers, 0, 1, 2, ...

These orbitals, e.g., 1s, 2s, 2p, 3d, 4f, represent different energy levels within the principal quantum number of an atom which can be occupied by electrons. The electrons are referred to as 1s electron, 2p electrons and so on.

3. The **third quantum number, m,** have integral values from -**l**, - (**l**-1), - (**l**-2) ... 0 ... (**l**-2), (**l**-1), **l**. The value **m** gives the number of alternative orientations of an orbital in space, e.g., if **l**=0 (which represent s-orbital) m = O, this is because an s-orbital is spherically symmetrical and therefore, there is no possible alternative orientation. If **l** =1 (p-orbital), then m =-1, 0, and 1 (= 3 values.) There are 3 alternative orientations of a p- orbital in space (Fig 8).

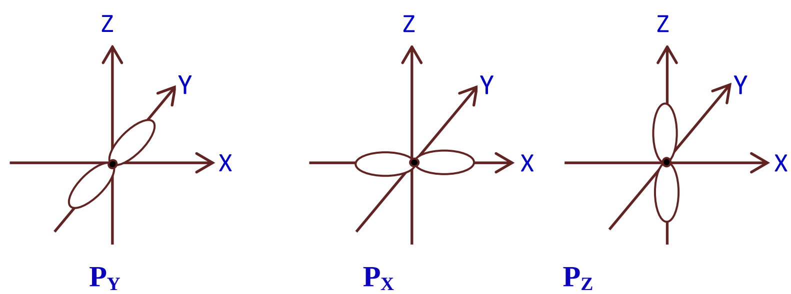


Fig. 8. Alternative orientation of p-orbital in space

4. The **spin quantum number, s,** has values -½ and +½. The value of **s** describes the spin of an election about its own axis, i.e., it can spin in only two ways: clockwise or anti-clockwise.

Definitions

An **atomic orbital** is a volume of space around the nucleus of an atom where there is a high probability of finding an electron. Fig. 8 below gives the shapes of the common orbital while Fig. 9 shows the box diagram representations of the sub-level in each of the orbital

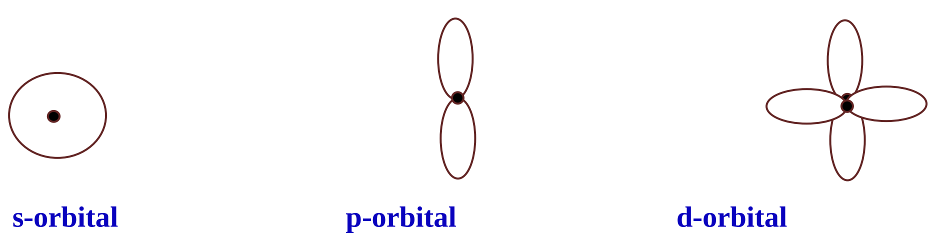


Fig. 8. Shapes of atomic orbital

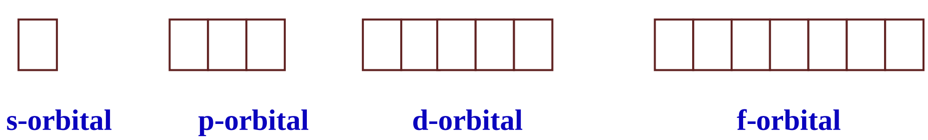


Fig. 9. The box diagram representations of the sub-level in each of the orbital

Each box in each sub-level is called an orbital and can hold a maximum of 2 electrons of opposite spins. Two electrons in the same orbital are said to be paired, i.e. $\boxed{\uparrow\downarrow}$

The maximum number of electrons that can be accommodated in each sub level is represented in table 3.

Table 3. The maximum number of electrons the can be accommodated in a given sub-level

Sub level	s	p	d	f
Maximum number of electron	2	6	10	14

The maximum number of electrons accommodated in each of the first four quantum shells is given in table 4

Table 4. The maximum number of electrons in each principal quantum shell

Quantum shell	**Maximum number of electrons**
K	2 electrons in 1s-orbital
L	8 (2 electrons in 2s- orbital and 6 electrons in 2p-orbital)
M	18 (2 electrons in 3s-, 6 electrons in 4p- and 10 electrons in 3d- orbital)
N	32 (2 electrons in 4s-, 6 electrons in 4p-, 10 electrons in 4d- and 14 electrons in 4f orbital.)

Generally the maximum number of electrons in a given quantum shell is $2n^2$: where n is the principal quantum number.

The relative energies of the sub shells (orbital) increase in the following order:

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d.

This order of energy levels can easily be worked out from Fig. 10.

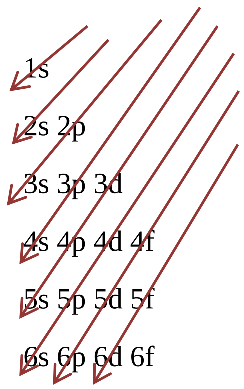


Fig. 10. Method of determining the order of energy level

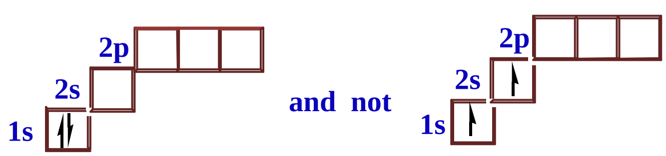
The rules governing the filling up orbital

The electrons are assigned to an atom by placing them into the various atomic orbital according to three rules.

1. **The Pauli Exclusion Principle**. No two electrons may have the same set of four quantum numbers. Where two electrons occupy the same orbital, they must have opposite spins: s = +½ for one electron and s = -½ for the second electron. Because the spin quantum number, **s,** can take only one of the two values; an orbital can house at most two electrons.

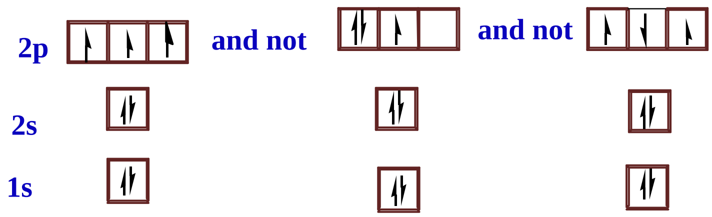
2. **The Aufban principle.** The electronic configurations are built up from the bottom, using the lowest energy orbital first.

e.g. Helium (2 electrons)



3. **Hund's rule**. Where orbitals are available in degenerate set, maximum spin multiplicity is preserved; that is, electrons are not paired until each orbital in the degenerate set has been half filled. The spins of any unpaired electrons are parallel.

e.g. Nitrogen (7 electrons)



Electronic configurations

The electronic configuration of an element is constructed, by assuming that electrons occupy the lowest possible energy level, being available, and the number of electrons in anyone level being determined by the four quantum numbers and the Pauli Exclusion Principle.

For instance, the electronic configuration of helium in the ground state is written as $1s^2$. Where the first value being the value of principal quantum number n, the letter **s** denoting the subsidiary quantum number **l** = 0 and the superscript indicating that there are two electrons in the same orbital with opposite spin. The notation  is usually used to show electron pair in an orbital with opposite spin.

Table 5. The electronic configurations of the first ten elements

Elements	Electronic configuration
H (1e$^-$)	$1s^1$
He (2e$^-$)	$1s^2$
Li (3e$^-$)	$1s^2\ 2s^1$
Be (4e$^-$)	$1s^2\ 2s^2$
B (5e$^-$)	$1s^2\ 2s^2\ 2p^1$
C (6e$^-$)	$1s^2\ 2s^2\ 2p^2$
N (7e$^-$)	$1s^2\ 2s^2\ 2p^3$
O (8e$^-$)	$1s^2\ 2s^2\ 2p^4$
F (9e$^-$)	$1s^2\ 2s^2\ 2p^5$
Ne (10e$^-$)	$1s^2\ 2s^2\ 2p^6$

Trial 2.2

1. *The electronic configuration of element X is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$*

   (a) *State with reason the period and group to which X belongs (2marks)*
   X is in period 4 because it in four energy levels/quantum shell and in group 1 because it contains only one electron in the outermost shell

   (b) *What is the valence of element X? Explain your answers (1 ½ marks)*
   X has a valence of +1 because it loses one electron to form a stable configuration

   (c) *Write electron configuration of atoms with the following atomic number; 24, 13, and 56*
   24 = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^4 4s^2$
   13 = $1s^2 2s^2 2p^6 3s23p^3$
   56 = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^2 4d^{10} 5p^6 6s^2$.

2. Write electron configuration of element with the following atomic numbers; in each case state the period and group in which each element belongs

   (i) 82

   (ii) 16

   (iii) 38

   (iv) 53

Nuclear reaction

This is a reaction where rearrangement of protons and neutrons in the nucleus of an atom take place and new element is formed.

Types of radiations

There are three types of radiations given by radioactive substances. They all cause certain substances, such as zinc sulphide, to luminesce, and all ionize gases through which they pass. They differ in their response to an electric field in the manner shown in figure below:

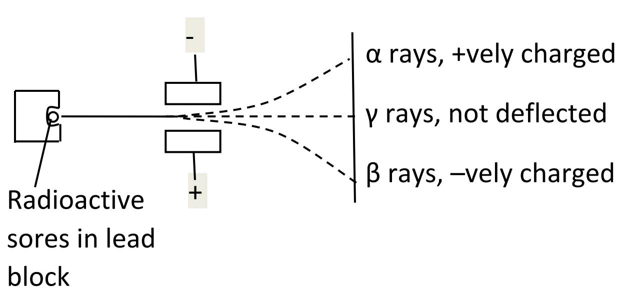


Fig. 11: Response of different types of radiation to electric field

γ-rays

- These uncharged rays are similar to X-rays
- they have high penetrating power; being able to pass through 0.1m of metal.
- have negligible weight

- are un deflected by electric field
- ionize gases they pass through

α-rays

- positively charged helium ions
- ionize gases they pass through
- deflected towards negative electric field
- have low penetrating power

β-rays

- negatively charged
- deflected toward positive electric field
- have medium penetrating power
- ionize gases they pass through

Balancing nuclear equations

The sum of protons and the mass number on the either side of the equation should be equal. deficits are balanced with either α-particle (^4_2He), β-particle ($_{-1}^{\ 0}e$, or $_{-1}^{\ 0}\beta$) or positron($_{+1}^{\ 0}e$) or neutron (^1_0n).

Example 2

The nucleus of element X reacts with an alpha particle according to the following equation

$^m_nX + ^4_2He \longrightarrow ^{45}_{21}Sn + ^1_0n$

Determine the values of m and n.

Solution

m+4 = 45 +1

m = 42

n + 2 = 21 + 0

n = 19

∴ X is potassium

Radioactivity

This is the spontaneous disintegration of unstable atoms with emission of particles like alpha, beta particles and gamma radiations.

The decay law

It states that the rate of disintegration of the nuclei in a given time is propotional to the number of atoms present

Rate of decay,R = $-\lambda\frac{dN}{dt}$

where N- number of atoms present, t = time, λ is a constant and negative because the number of atoms are reducing

The decay law can also be xpressed as

$N = N_0e^{-\lambda t}$

where No is the intial number of disintegrating atoms.

The decay constant is the fractional number of atoms that are disintegrating per second

Half-life ($t_{½}$) is the time taken for the number of atoms in a radioactive element to reduce to half the original value.

From $N = N_0e^{-\lambda t}$

$\ln \frac{N_0}{N} = \lambda t$

At t = t ½ ; $N = \frac{N_0}{2}$

⇨ $\ln \frac{N_0}{N/2} = \lambda t½$

⇨ $t½ = \frac{\ln 2}{\lambda} = \frac{0.693}{\lambda}$

Activity is the rate of dsintegration of a radioactive substance = λN

Example 3

A sample of radioactive material initially contains 10^{18} atoms. If the half-life of the material is 2 days, calculate the

(i) number of atoms remaining after 5days

$\lambda = \frac{0.693}{t_{\frac{1}{2}}} = \frac{0.693}{2} = 0.3465s^{-1}$

$N = 10^{18}e^{-0.3467 \, x \, 5} = 1.7684 \times 10^{17}$

(ii) percentage that decayed after 5days

Number of decayed atoms = $N_0 - N$

Percentage decayed = $\frac{N_0 - N}{N_0} \, x \, 100\% = \frac{10^{18} - 1.7684 \, x \, 10^{17}}{10^{18}}$x100% = 82.32%

(iii) activity of the sample after 5days

Activity, A = λN = 0.3465 x 1.7684 x 10^{17} = 6.127506 cx 10^{16}

Trial 2.3

1. Polonium $^{216}_{84}Po$ undergoes radioactive to give element Y according to the following equation $^{216}_{84}Po \rightarrow Y + \alpha$

   (a) Calculate

   (i) atomic number of Y [82]

   (ii) the mass number of Y [212]

   (b) Y decays further to form Z as shown by the equation below

   $Y \rightarrow Z + \beta$

   Calculate

(i) the atomic number of Z [83

(ii) the mass number of Z [212]

2. The following equation shows part off the radioactive decay of Thorium.

$^{234}_{90}Th \rightarrow {}^{234}_{91}Pa$ + ___ $\rightarrow$ X +α

(i) Name the particle emitted in the first stage of the reaction [β]

(ii) State the atomic number and the atomic mass of X [89, 230] (1mk)

3. (a) State three properties of beta particles

(b) complete the following nuclear transformations

(i) $^{239}_{94}Pu + {}^{1}_{0}n \rightarrow {}^{86}_{34}Se + {}^{150}_{60}Nd$ + _____ $[\mathbf{4{}^{1}_{0}n}]$

(ii) $^{235}_{96}Cm + {}^{4}_{2}He \rightarrow$ ______ $[\mathbf{{}^{239}_{98}Cf}]$

(iii) $^{235}_{92}U + {}^{1}_{0}\mathrm{n} \rightarrow {}^{148}_{57}La$ + ____ + $3{}^{1}_{0}n$ $[\mathbf{{}^{85}_{35}Br}]$

(c) Francium isotope ($^{223}_{87}Fr$) emits beta particles. the rate of emission reduces from 14.0 to 7.5 counter in 80 second. Calculate the half-life of isotopes. [88.865s]

4. Complete the following equations

(a) $^{234}_{91}\mathrm{Pa} \rightarrow {}^{234}_{92}\mathrm{U}$ + -------- $[\mathbf{{}^{0}_{-1}\beta}]$

(b) $^{6}_{3}\mathrm{Li} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{3}_{1}\mathrm{H}$ +--------- $[\mathbf{{}^{4}_{2}He}]$

(c) $^{107}_{47}\mathrm{Ag} + {}^{1}_{0}\mathrm{n} \rightarrow$……………… $[{}^{108}_{47}Ag]$

5. (a) Complete the following equations for the decay of bismuth.

$^{214}_{83}\mathrm{Bi} \rightarrow {}^{0}_{-1}\mathrm{e}$ + -------- $[\mathbf{{}^{214}_{84}po}]$

(b) The half-life of bismuth is 19.7minutes. Determine the time taken for 43% by mass of bismuth to decay. [16s]

hank you
Dr. Bbosa Science