Nuclear Chemistry

"The branch of chemistry which deals with the study of composition of atomic nucleus and the nuclear transformations"

The discussion of nuclear science with special emphasis on its chemical aspects is termed **Nuclear chemistry.** It has become a very important and fascinating branch of science due to the tremendous amount of energy liberated during nuclear reaction, which led to the discovery of atom bomb, hydrogen bomb, etc. The controlled release of nuclear energy promises to lead us into a new era, in which an unlimited storehouse of energy is available to man. It is rightly said that we are now living in the nuclear age and the future of humanity is at the mercy of the nuclear scientists. They can annihilate the whole world within a few minutes. They are the hope of human happiness and prosperity as well. In this chapter we shall deal broadly with the various aspects of nuclear chemistry

Nucleus.

Nucleus is found to be a source of tremendous amount of energy, which has been utilised for the destructive as well as constructive purposes. Hence the study of nucleus of an atom has become so important that it is given a separate branch of chemistry under the heading of nuclear chemistry. According to an earlier hypothesis, the nucleus is considered as being composed of two building blocks, **proton's** and **neutron's**, which are collectively called **nucleons**.

(1) **Nuclear forces :** Since the radius of nucleus is very small $\approx 10^{-15} m$, two protons lying in the nucleus are found to repel each other with an electrostatic force of about 6 tonnes. *The forces, which hold the nucleons together means stronger proton – proton, neutron – neutron and even proton – neutron attractive forces, exist in the nucleus. These attractive forces are called nuclear forces.* Unlike electrostatic forces which operate over long ranges, but the nuclear forces operate only within small distance of about $1 \times 10^{-15} m$ or 1 *fermi* (1 *fermi* = $10^{-13} cm$) and drops rapidly to zero at a distance of $1 \times 10^{-13} cm$. Hence these are referred to as short range forces. Nuclear forces are nearly 10^{21} times stronger than electrostatic forces.

(2) **Nuclear stability :** Nucleus of an atom contains all the protons and neutrons in it while all electrons are in the outer sphere. Nuclides can be grouped on the basis of nuclear stability, *i.e.* stable and unstable nucleus. The most acceptable theory about the atomic nuclear stability is based upon the fact that the observed atomic mass of all known isotopes (*except hydrogen*) is always less from the sum of the weights of protons and neutrons **(nucleons)** present in it. Other less important (or unusual) fundamental particles of the nucleus are electron, antiproton, positron, neutrino, photon, graviton, meson and γ - particles are considered as created by stresses in which energy is converted into mass or vice versa, *e.g.* an electron (β- particle) from a radio active nucleus may be regarded as derived from a neutron in the following way.

$$\textbf{Neutron} \longrightarrow \textbf{Proton + Electron}$$

Similarly, photons are produced from internal stresses within the nucleus.

A list of elementary particles is given below :

Group	Name of particle	Symbol	*Anti – particle symbol*	*Mass*	*Spin*	*Charge*
LEPTONS	Photon	hν	–	0	1	0
LEPTONS	Electron	e^-	e^+	1	1/2	–1
LEPTONS	Neutrino	Ve	$\bar{V}e$	0	1/2	0
LEPTONS	Muon	μ^-	μ^+	207	1/2	–1
LEPTONS	Muon –neutrino	V_μ	$\bar{V}_\mu$	0	1/2	0
LEPTONS	Tauon	τ^-	τ^+	3500	1/2	–1
HARDONS – MESONS		π^0		264	0	0
HARDONS – MESONS	Pions	π^+	π^-	273	0	+1
HARDONS – MESONS	Kaons	k^+	k^-	966	0	+1
HARDONS – MESONS		k^0	$\bar{k}^0$	974	0	0
HARDONS	Etameson	n^0	–	1074	0	0
HARDONS	Proton	p	p^-	1836.6	1/2	+1
HARDONS	Neutron	n	n^-	1836.6	1/2	0
HARDONS – BARYONS	Lambda hyperon	λ^0	$\bar{\lambda}^0$	2183		
HARDONS – BARYONS		Σ^+	$\bar{\Sigma}^+$	2328	1/2	0
HARDONS – BARYONS	Sigma hyperons	Σ^0	$\bar{\Sigma}^0$	2334	1/2	+1
HARDONS – BARYONS		Σ^-	$\bar{\Sigma}^{-1}$	2343	1/2	0
HARDONS					1/2	–1
HARDONS	Xi hyperons	1.1, 1.1	–	2573	1/2	0
HARDONS	Omega hyperons	Ω^-	$\bar{\Omega}^-$	3273	3/2	–1

Some common important elementary particles are listed below :

Name	*Symbol*	*Mass*	*Charge*	*Discoverer*
Electron	e^-	$9.1\times10^{-31}kg$	$-1.602\times10^{-19}C$	J.J. Thomson (1896)
Proton	p	$1.673\times10^{-27}kg$	$+1.602\times10^{-19}C$	E. Goldstein (1886)
Neutron	n	$1.675\times10^{-27}kg$	Zero	J. Chadwick (1932)
Neutrino	V	$3.64\times10^{-32}kg$	Zero	Pauli
Mesons	μ	275 – 300 times mass of electron	$+ve, 0$ or $-ve$	Yukawa (1935)
Positron	e^+	$9.1\times10^{-31}kg$	$+ve$	Anderson (1932)

The stability of nucleus may be discussed in terms of any one of the following.

(i) **Nuclear Binding Energy and Mass defect :** The mass of hydrogen atom is equal to the sum of the masses of a proton and an electron. For other atoms, the atomic mass is less than the sum of the masses of protons, neutrons and electrons present. This difference in mass termed as, *mass defect,* is a measure of the *binding energy* of protons and neutrons in the nucleus. The mass energy relationship postulated by Einstein is expressed as :

$\boxed{\Delta E = \Delta mc^2}$, Where ΔE is the energy liberated, Δm the loss of mass and c is the speed of light.

Consider the helium nucleus, which contains 2 protons and 2 neutrons; the mass of helium nucleus on $^{12}C = 12m_u$, scale is $4.0017m_u$. The masses of individual isolated proton and neutron are 1.0073 and 1.0087 m_u respectively. The total mass of 2 protons and 2 neutrons is $(2 \times 1.0073) + (2 \times 1.0087) = 4.0320m_u$. The loss in mass or mass defect for helium nucleus is, $4.0320m_u - 4.0017m_u = 0.0303m_u$

$\because \quad 1m_u = 1.66057 \times 10^{-27} kg$ and $c = 2.998 \times 10^8 ms^{-1}$

$$\Delta E = 0.0303 \times 1.66057 \times 10^{-27} \times 6.02 \times 10^{23} \times (2.998 \times 10^8)^2 \; kg\, m^2 s^{-2} mol^{-1} \; = 2.727 \times 10^{12} \; J\, mol^{-1}$$

Thus, the molar nuclear binding energy of helium nucleus, ${}_2He^4$, is $2.73 \times 10^{12} J\, mol^{-1}$. Binding energy of a nucleus is generally quoted as energy in million electron volts (*MeV*) per nucleon. One million electron volts are equivalent to $9.6 \times 10^{10} J\, mol^{-1}$. Thus, the formation of helium nucleus results in the release of $2.7 \times 10^{12} / 9.6 \times 10^{10}$ *MeV* = 28 *MeV* (approximately). In comparing the binding energies of different nuclei, it is more useful to consider the binding energy per nucleon. For example, helium nucleus contains 4 nucleons (2 protons and 2 neutrons), the binding energy per nucleon in this case is 28/4 = 7 *MeV*.

Binding energies of the nuclei of other atoms can be calculated in a similar manner. When we plotted binding energies of the nuclei of atoms against their respective mass number. Three features may be noted. First, nuclei with mass number around 60 have the highest binding energy per nucleon. Second, species of mass numbers 4, 12, and 16 have high binding energy per nucleon implying that the nuclei ^4He, ^{12}C and ^{16}O, are particularly stable. Third the binding energy per nucleon decreases appreciably above mass number 100. The form of relationship between binding energy per nucleon and mass number indicates that heavy nuclei would release mass (and therefore energy) on division (or fission) into two nuclei of medium mass and that the light nuclei would release mass (and therefore energy) on fusion to form heavier nuclei. These processes called fission and fusion are described later in this Unit.

- The **average binding energy** for most of the nuclei is in the vicinity of 8 MeV. Nuclei having binding energy per nucleon very near to 8 MeV are more or less stable.
- Iron has the maximum average binding energy (8.79 MeV) and thus its nucleus is thermodynamically most stable.
- The isotopes with intermediate mass numbers 40 to 100 are most stable. The elements with Low Mass numbers or High Mass numbers tend to become stable by acquiring intermediate mass number. Evidently, nuclei of lighter elements combine together to form a heavier nucleus of intermediate mass number **(nuclear fusion)**; while the nuclei of heavy elements split into two lighter nuclei of intermediate mass numbers **(nuclear fission).** In either case, energy is released and hence the stability is enhanced.

Note : ✸ Relation between different units of energy $1cal = 4.2J$; $1\,J = 10^7\, ergs$; $1eV = 1.622 \times 10^{-19} J$

(ii) **Relative stability of isotopes and binding energy:** Value of binding energy predicts the relative stability of the different isotopes of an element. If the value of binding energy is negative, the product nucleus or nuclei will be less stable than the reactant nucleus. Thus the relative stability of the different isotopes of an element can be predicted by the values of binding energy for each successive addition of one neutron to the nucleus.

$${}_2He^3 + {}_0n^1 \longrightarrow {}_2He^4 + 20.5MeV\,; \quad {}_2He^4 + {}_0n^1 \longrightarrow {}_2He^5 - 0.8MeV$$

Therefore, ${}_2He^4$ is more stable than ${}_2He^3$ and ${}_2He^5$.

(iii) **Packing fraction:** The difference of actual isotopic mass and the mass number in terms of packing fraction is defined as:

$$\text{Packing fraction} = \frac{\text{Actual isotopic mass} - \text{Mass number}}{\text{Mass number}} \times 10^4$$

The value of packing fraction depends upon the manner of packing of the nucleons with in the nucleus. Its value can be negative, positive or even zero.

Note : ✹ Actual isotopic mass is not a whole number whereas, mass number is a whole number.

(a) **Zero packing fraction:** $Carbon^{12}$ has zero packing fraction because it is taken as a reference on the atomic scale and its actual isotopic mass (12) is equal to its mass number (12).

(b) **Negative packing fraction :** Negative value of the packing fraction means that the actual isotopic mass is less to the mass number. This term indicates that some mass has been transformed into energy (binding energy) during formation of nucleus. Such nuclei are, therefore more stable.

(c) **Positive packing fraction :** Positive packing fraction should imply the opposite, *i.e.,* the nuclei of such isotopes should be unstable. However, this generalisation is not strictly correct especially for elements of Low Mass numbers. For these elements, though packing fraction is positive, yet they are stable. This is explained on the basis that the actual masses of protons and neutrons (of which the nuclei are composed) are slightly greater than unity.

In general, lower the packing fraction, greater is the binding energy per nucleon and hence greater is the stability *the relatively low packing fraction of He, C and O implies their exceptional stability packing fraction is least for Fe (negative) and highest for H (+78).*

(iv) **Meson theory of nuclear forces :** Neutron is found to play a leading role in binding the nuclear particles. It has been established that neutron proton attractions are stronger than the proton-proton or neutron – neutron attraction. This is evident by the fact that the deutron, $_1H^2$ having one proton and one neutron, is quite stable.

Yukawa in 1935, put forward a postulate that neutrons and protons are held together by very rapid exchange of nuclear particles called *Pi-mesons (π-mesons have mass equal to 275 times of the mass of an electron and a charge equal to +1, 0 or –1. There are designated as π^+ π^0 and π^- respectively*). The nuclear force which is used in rapid exchange of *Pi*-mesons between nucleons are also called **exchange forces.**

- The binding forces between unlike nucleons (*p* and *n*) are explained by the oscillation of a charged π-meson (π^+ or π^-)

$$\text{(a)}\quad p_1 + n_2 \rightleftharpoons n_1 + \pi^+ + n_2 \rightleftharpoons n_1 + p_2$$
$$\text{(b)}\quad p_1 + n_2 \rightleftharpoons n_1 + \pi^- + p_2 \rightleftharpoons n_1 + p_2$$

- Binding forces between like nucleons (*p* - *p* or *n* - *n*) result from the exchange of neutral mesons (π^0) as represented below.

$$\text{(a)}\quad p_1 \rightleftharpoons p_2 + \pi^0 \text{ or } p_1 + \pi^0 \rightleftharpoons p_2$$
$$\text{(b)}\quad n_1 \rightleftharpoons n_2 + \pi^0 \text{ or } n_1 + \pi^0 \rightleftharpoons n_2$$

(v) **Nuclear shell model :** According to this theory, nucleus of atom, like extra-nuclear electrons, also has definite energy levels (shells). The shell structure is supported by the existence of periodicity in the nuclear properties. For example, elements with even number of protons and neutrons are more abundant, more stable and richer in isotopes. Nuclides with odd number of protons and neutrons are least abundant in nature (only 5 are known ${}_1H^2$, ${}_5B^{10}$, ${}_7N^{14}$ and ${}_{73}Ta^{180}$).

Thus elements have a tendency to have even number of both protons and neutrons. This suggests that like electrons, nucleon particles in the nucleus are paired. Magnetic fields of the two paired protons spinning in opposite direction cancel each other and develop attractive forces, which are sufficient to stabilize the nucleus. Further nuclei with 2, 8, 20, 28, 50, 82 or 126 protons or neutrons have been found to be particularly stable with a large number of isotopes. These numbers, commonly known as **Magic numbers** are defined as the number of nucleons required for completion of the energy levels of the nucleus. Nucleons are arranged in shells as two protons or two neutrons (with paired spins) just like electrons arranged in the extra-nuclear part. Thus the following nuclei ${}_2He^4$, ${}_8O^{16}$, ${}_{20}Ca^{40}$ and ${}_{82}Pb^{208}$ containing protons 2, 8, 20 and 82 respectively (all magic numbers) and neutrons 2, 8, 20 and 126 respectively (all magic numbers) are the most stable.

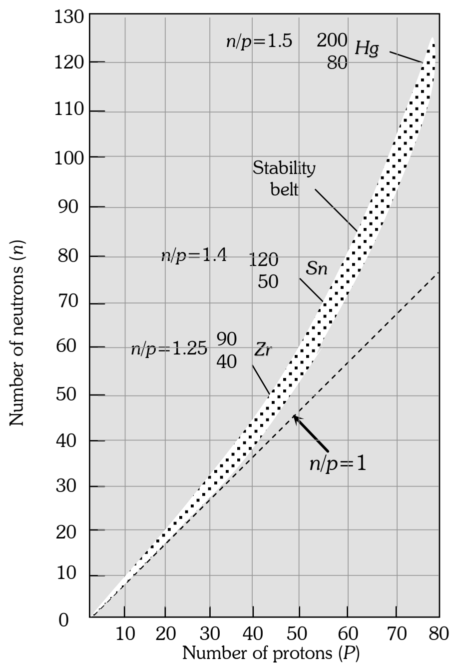
Magic numbers for **protons** : 2, 8, 20, 28, 50, 82, 114

Magic numbers for **neutrons** : 2, 8, 20, 28, 50, 126, 184, 196

When both the number of protons and number of neutrons are magic numbers, the nucleus is very stable. That is why most of the radioactive disintegration series terminate into stable isotope of lead (magic number for proton = 82, magic number for neutron = 126). Nuclei with nucleons just above the magic numbers are less stable and hence these may emit some particles to attain magic numbers.

(vi) **Nuclear fluid theory :** According to this theory the nucleus is considered to resemble a liquid drop. Nucleons are believed to be present in the nucleus as nuclear fluid of very high density equal to 130 trillion tonnes/m^3, which is about 100 trillion times the density of water. The density is uniform and does not vary from atom to atom. Along with its almost unbelievable high density nuclear fluid possesses a correspondingly high surface tension ($= 9.3 \times 10^{19}\, Nm^{-1}$, *i.e.,* 1.24×10^{18} times the surface tension of water). A nuclear film attached to a wire one centimetre long would support the mass of one billion tonnes. This force of nuclear surface tension is, in fact, responsible for keeping the nucleons bound together against the forces of repulsion. This is known as the **Nuclear fluid theory** of the stability of the nuclei. Thus according to this theory the nucleons are free to move with in the nucleus whereas according to the nuclear shell structure theory the nucleons exist in definite energy levels.

(vii) **Neutron-proton ratio and nuclear stability or causes of radioactivity :** The nuclear stability is found to be related to the neutron/proton (n/p) ratio. If for different elements the number of neutrons is plotted against the number of protons, it is found that the elements with stable nuclei (non-radioactive elements) lie within a region (belt) known as **zone** or **belt of stability**.



Nuclei lying above or below the stability belt are unstable

(a) For elements with low atomic number (less than 20), *n/p* ratio is 1, *i.e.,* the number of protons is equal to the number of neutrons. Remember that *n/p* ratio of $_1H^1$ is zero as it has no neutron. Nuclide with highest *n/p* ratio is $_1H^3$ (*n/p* = 2.0)

(b) With the increase in atomic number although the number of protons increases but the number of neutrons increases much more than the number of protons with the result the *n/p* ratio goes on increasing from 1 till it becomes nearly equal to 1.5 at the upper end of the belt.

(c) When the *n/p* ratio exceeds 1.52 as in elements with atomic number 84 or higher, the element becomes radioactive and undergoes disintegration spontaneously. Note that these elements lie outside the zone of stability.

The way an unstable nucleus disintegrates is decided by its position with respect to the actual *n/p* plot of stable nuclei (the zone of stability)

- ***Neutrons to proton (n/p) ratio too high.*** If the *n/p* ratio is too high, *i.e.,* when the nucleus contains too many neutrons, it falls above the zone of stability. The isotope would be unstable and would tend to come within the stability zone by the emission of a β-ray (electron). Electron, is produced in the nucleus probably by the following type of decay of a neutron.

$$\boxed{_0n^1 \rightarrow _1H^1 + _{-1}e^0} \text{ (Beta particle)}$$

The electron thus produced is emitted as a β-particle and thus the neutron decay ultimately increases the number of protons, with the result the *n/p* ratio decreases and comes to the stable belt. Consider the example of C^{12} and C^{14}. In C^{12}, the *n/p* ratio (6/6) is 1, hence its nucleus is quite stable. On the other hand, in C^{14}, the *n/p* ratio (8/6) is 1.3, hence it should be unstable. In practice also it is found to be so and C^{14} decays in he following way to give N^{14} (*n/p* ratio = 1)

$$\underset{\left(\frac{n}{p}=\frac{8}{6}=1.33\right)}{_6C^{14}} \longrightarrow \underset{\left(\frac{n}{p}=\frac{7}{7}=1.0\right)}{_7N^{14}} + _{-1}e^0 \; ; \quad \underset{\left(\frac{n}{p}=\frac{146}{92}=1.587\right)}{_{92}U^{238}} \longrightarrow \underset{\left(\frac{n}{p}=\frac{144}{90}=1.6\right)}{_{90}Th^{234}} + _2He^4$$

- **Neutron to proton ratio (*n/p*) too low**, (*i.e.,* when the nucleus contains excess protons) : There are no naturally occurring nuclides with *n/p* ratio less than 1, however there are many artificially nuclides in such cases, the nucleus lies below the zone of stability, it would again be unstable and would tend to come within the zone of stability by losing a positron.

$$\underset{\left(\frac{n}{p}=\frac{5}{6}=0.83\right)}{_6C^{11}} \longrightarrow \underset{\left(\frac{n}{p}=\frac{6}{5}=1.2\right)}{_5B^{11}} + _{+1}e^0; \quad \underset{\left(\frac{n}{p}=\frac{5}{6}=0.83\right)}{_7N^{13}} \longrightarrow \underset{\left(\frac{n}{p}=\frac{7}{6}=1.16\right)}{_6C^{13}} + _{+1}e^0$$

Such nuclides can increase *n/p* ratio by adopting any one of the following three ways :

By emission of an alpha particle : $\underset{\left(\frac{n}{p}=\frac{146}{92}=1.58\right)}{_{92}U^{238}} \rightarrow \underset{\left(\frac{n}{p}=\frac{144}{90}=1.60\right)}{_{90}Th^{234}} + _2He^4$

By emission of a positron : $\underset{\left(\frac{n}{p}=\frac{6}{7}\right)}{^{13}_7N} \rightarrow \underset{\left(\frac{n}{p}=\frac{7}{6}\right)}{^{13}_6C} + _{+1}e^0$

By K-electron capture : $\underset{\left(\frac{n}{p}=\frac{115}{79}\right)}{^{194}_{79}Au} + _{-1}e^0 \rightarrow \underset{\left(\frac{n}{p}=\frac{116}{78}\right)}{^{194}_{78}Pt}$

α-emission is usually observed in natural radioactive isotopes while emission of positron or *K*-electron capture is observed in artificial radioactive isotopes. The unstable nuclei continue to emit α or β-particles. Until stable nuclei comes into existence.

(3) **Nuclear reactions :** In a chemical reaction, only electrons (extra-nuclear particle) of the atom take part while the nucleus of the atom remains unaffected. However, the reverse reactions (*i.e.*, where only nuclei of atoms take part in reactions) are also possible. *Such reactions in which nucleus of an atom itself undergoes spontaneous change or interact with other nuclei of lighter particles resulting new nuclei and one or more lighter particles are called nuclear reactions.*

(i) **Some characteristics of nuclear reactions :**

(a) *Nuclear reactions are written like a chemical reaction* : As in a chemical reaction, reactants in a nuclear reaction are written on the left hand side and products on the right hand side with an arrow in between them.

(b) *Mass number and atomic number of the elements are written in a nuclear reactions* : Mass number and atomic number of the element involved in a nuclear reaction are inserted as superscripts and subscripts respectively on the symbol of the element. For example ${}^{27}_{13}Al$ or Al^{27}_{13} or ${}_{13}Al^{27}$ stands for an atom of aluminum with mass number 27 and atomic number 13.

(c) *Mass number and atomic number are conserved* : In a nuclear reaction the total mass numbers and total atomic numbers are balanced on the two sides of the reaction (recall that in an ordinary reaction the total number of atoms of the various elements are balanced on the two sides)

(d) Energy involved in the nuclear reactions is indicated in the product as $+Q$ or $-Q$ of reactions accompanied by release or absorption of energy respectively.

(e) Important projectiles are α-particles $({}_2He^4)$, Proton $({}_1H^1$ or $p)$, deutron $({}_1H^2$ or ${}_1D^2)$, neutron $({}_0n^1)$, electron (β-particle or ${}_{-1}e^0$ or e^-) and positron $({}_{+1}e^0)$.

(f) *Representation of nuclear reactions* : For example, ${}_7N^{14} + {}_2He^4 \rightarrow {}_8O^{17} + {}_1H^1 + Q$. Some times a short hand notation is used, *e.g.,* the above reaction can be represented as below. ${}_7N^{14}(\alpha, p)_8O^{17}$

(ii) **Nuclear reactions Vs chemical reactions :**

(a) As per definition, chemical reactions depend upon the number of extranuclear electrons while nuclear reactions are independent upon the electrons but depend upon the nature of the nucleus.

(b) Chemical reactions involve some loss, gain or overlap of outer orbital electrons of the two-reactant atoms. On the other hand, nuclear reactions involve emission of some light particles (α, β, positron, etc.) from the nucleus of the atom to form another element.

(c) The chemical reactivity of the element is dependent on the nature of the bond present in the concerned compound. On the other hand, the nuclear reactivity of the element is independent of its state of chemical combination, *e.g.,* radium, whether present as such or in the form of its compound, shows similar radioactivity.

(d) The energy change occuring in nuclear reactions is very high as compared to that in chemical reactions. Again in chemical reactions the energy is expressed in *kcal* per mole while in nuclear reactions the energy is expressed in *MeV* per nucleus. Nuclear reactions, which liberate energy are called **exoergic reactions** and which absorb energy are called **endoergic.**

(*e*) A chemical reaction is balanced in terms of mass only while a nuclear reaction must be balanced in terms of both mass and energy. In endoergic reactions, the mass of products is more than the mass of reactants. While in exoergic reaction the mass of products is less than the mass of reactants.

(f) The chemical reactions are dependent on temperature and pressure while the nuclear reactions are independent of external conditions.

(iii) **Types of nuclear reactions :** Nuclear reactions may broadly be divided into two types :

(a) **Natural nuclear reactions :** In these reactions, nucleus of a single atom undergoes a spontaneous change itself.

(b) **Artificial nuclear reactions :** In these reactions, two nuclei of different elements are brought to interact artificially. Bombarding a relatively heavier nucleus (non-radioactive) with a lighter nucleus, viz. proton, deutron and helium, does this. Artificial nuclear reactions are divided as follows :

- **Projectile capture reactions :** The bombarding particle is absorbed with or without the emission of γ-radiations.

$$_{92}U^{238} + {}_0n^1 \rightarrow {}_{92}U^{239} + \gamma\,; \quad {}_{13}Al^{27} + {}_0n^1 \rightarrow {}_{13}Al^{28} + \gamma$$

- **Particle-particle reactions :** Majority of nuclear reactions come under this category. In addition to the product nucleus, an elementary particle is also emitted.

$$_{11}Na^{23} + {}_1H^1 \rightarrow {}_{12}Mg^{23} + {}_0n^1\,; \quad {}_{11}Na^{23} + {}_1H^2 \rightarrow {}_{11}Na^{24} + {}_1H^1$$

$$_{11}Na^{23} + {}_2He^4 \rightarrow {}_{12}Mg^{26} + {}_1H^1\,; \quad {}_7N^{14} + {}_0n^1 \rightarrow {}_6C^{14} + {}_1H^1$$

- **Spallation reactions :** High speed projectiles with energies approximately 40 *MeV* may chip fragments from a heavy nucleus, leaving a smaller nucleus.

$$_{29}Cu^{63} + {}_2He^4 + 400\,MeV \rightarrow {}_{17}Cl^{37} + 14\,{}_1H^1 + 16\,{}_0n^1$$

- **Fission reactions :** A reaction in which a heavy nucleus is broken down into two or more medium heavy fragments. The process is usually accompanied with emission of neutrons and large amount of energy.

$$_{92}U^{235} + {}_0n^1 \rightarrow {}_{56}Ba^{141} + {}_{36}Kr^{92} + 3\,{}_0n^1 + 200\,MeV$$

- **Fusion reactions :** Light nuclei fuse together to reproduce comparatively heavier nuclei. A fusion reactions is the source of tremendous amount of energy.

$$_1H^2 + {}_1H^3 \rightarrow {}_2He^4 + {}_0n^1 + 17.6\,MeV$$

Example : 1 Sulphur-35 (34.96903 *amu*) emits a β-particle but no γ-ray. The product is chlorine-35 (34.96885 *amu*). The maximum energy emitted by the β-particle is **[CBSE 1999]**

(a) 16.758 *MeV* (b) 1.6758 *MeV* (c) 0.16758 *MeV* (d) 0.016758 *MeV*

Solution: (c) The mass converted into energy $= 34.96903\,amu - 34.96885\,amu = 1.8 \times 10^{-4}\,amu$ $(1\,amu = 931.5\,MeV)$

$\because$ Energy produced $= 1.8 \times 10^{-4} \times 931.5 = 0.16758\,MeV$

Example : 2 If the atomic masses of lithium, helium and proton are 7.01823 *amu*, 4.00387 *amu* and 1.00815 *amu* respectively, calculate the energy that will be evolved in the reaction. $Li^7 + H^1 \rightarrow 2He^4 + \text{energy}$ **.**

(Given that 1 *amu* = 931 *MeV*)

(a) 17.3 *MeV* (b) 17.8 *MeV* (c) 17.2 *MeV* (d) 17.0 *MeV*

Solution: (a) Total mass of the reacting species $(Li^7 \text{ and } H^1) = 7.01823 + 1.00815 = 8.02638$ *amu*

The mass of the resulting species $(2He^4) = 2 \times 4.00387 = 8.00774$ *amu*

Mass of reacting species converted into energy, *i.e.*, $\Delta m = 8.02638 - 8.00774 = 0.01864$ *amu*

$\therefore$ Energy evolved in the reaction $= 0.01864 \times 931.5 = 17.363\, MeV.$

Example : 3 Calculate the mass defect and binding energy per nucleon for $_{27}Co^{59}$. [The mass of $Co^{59} = 58.95\, amu$, mass of hydrogen atom = 1.008142 *amu* and mass of neutron = 1.008982 *amu*].

(a) 8.77 *MeV* (b) 8.25 *MeV* (c) 9.01 *MeV* (d) 8.00 *MeV*

Solution: (a) Number of protons in $_{27}Co^{59} = 27$

$\therefore$ Number of neutrons = 59 – 27 = 32

$\Delta m = (1.008142 \times 27 + 1.008982 \times 32) - 58.95 = 0.556438\, amu$

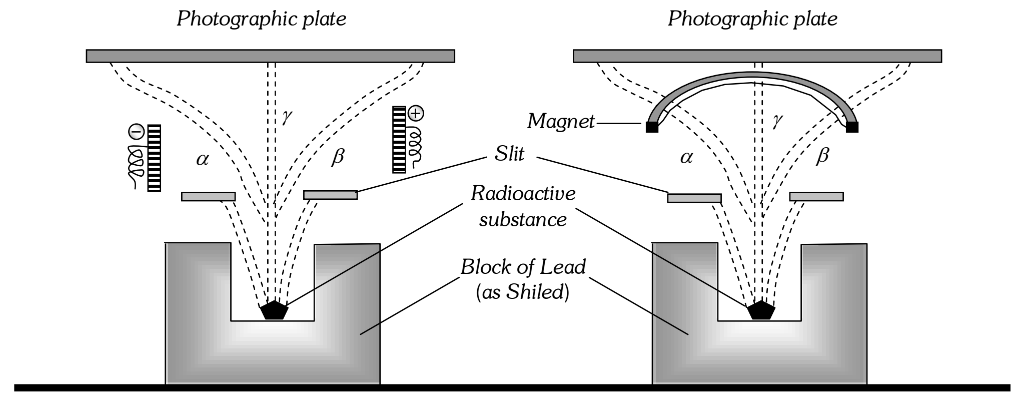
The binding energy (E_B) per nucleon $= \dfrac{\Delta m \times 931}{\text{Mass of cobalt}} = \dfrac{0.556438 \times 931}{59} MeV = 8.77\, MeV$

Radioactivity.

"Radioactivity is a process in which nuclei of certain elements undergo spontaneous disintegration without excitation by any external means."

- **Henry Becquerel (1891)** observed the spontaneous emission of invisible, penetrating rays from *potassium uranyl sulphate* $\boldsymbol{K_2UO_2(SO_4)_2}$, which influenced photographic plate in dark and were able to produce luminosity in substances like *ZnS*.
- Later on, **Madam Marie Curie and her husband P. Curie** named this phenomenon of spontaneous emission of penetrating rays as, **Radioactivity**. They also pointed out that radioactivity is a characteristic property of an unstable or excited nucleus, *i.e.*, a nuclear property is **independent** of all the external conditions such as pressure, temperature, nature of other atoms associated with unstable atom but **depends** upon the amount of unstable atom.
- **Curies** also discovered a new radioactive element **Radium** from *pitchblende* (an ore of *U i.e.* U_3O_8) which is about 3 million times more radioactive than uranium. Now a days about 42 radioactive elements are known.
- The elements whose atoms disintegrate and emit radiations are called radioactive elements.
- Radioactivity can be detected and measured by a number of devices like ionisation chamber, Geiger Muller counter, proportional counter, flow counter, end window counter, scintillation counter, Wilson cloud chamber, electroscope, etc. The proper device depends upon the nature of the radioactive substance and the type of radiation emitted. *GM counter and proportional counter are suitable for solids and liquids, ionisation chamber is most suitable for gases.*
- Lightest radioactive isotope is tritium $(_1H^3)$; other lighter radioactive nuclides are ^{14}C, ^{40}K and ^{99}Tc.

(1) **Nature of radioactive emissions :** The nature of the radiations emitted from a radioactive substance was investigated by **Rutherford (1904)** by applying electric and magnetic fields to the radiation as shown in figure.



Study of the nature of radiations emitted form a radioactive substance

It is observed that on applying the field, the rays emitted from the radioactive substances are separated into three types, called α, β, and γ-rays. The α-rays are deflected in a direction which shows that they carry positive charge; the β-rays are deflected in the opposite direction showing that they carry negative charge and the γ-rays are not deflected at all showing that they carry no charge.

(2) **Characteristics of radioactive rays :** Radioactive rays are characterised by the following properties :

(i) They blacken photographic plates.

(ii) They pass through thin metal foils.

(iii) They produce ionization in gases through which they passes.

(iv) They produce luminescence in zinc sulphide, barium platinocyanide, calcium tungstate, etc.

Radioactive radiations are composed of three important rays, namely α, β and $\gamma-$rays which differ very much in their nature and properties, *e.g.* penetrating power, ionising power and effect on photographic plates. Remember that γ-rays are not produced simultaneously with α and β-rays but are produced subsequently.

Comparison of α, β and γ-rays

α-Particle or α-Ray	β-Particle or β-Ray	γ-Ray
(1) Charge and mass : It carries 2 units positive charge and 4 unit mass.	**It carries 1 unit negative charge and no mass.**	**These are electromagnetic rays with very short wavelength (app. 0.05 Å)**
(2) Nature : It is represented as helium nucleus or helium ions $_2He^4$ **or** He^{++}.	**It is represented as electron** $_{-1}e^0$.	**It is represented as** $_0\gamma^0$
(3) Action of magnetic field : These are deflected towards the cathode.	**These are deflected to anode.**	**These are not deflected.**
(4) Velocity : $2\times10^9\ cm/s$ **or** $2\times10^7\ m/sec$ **(1/10th to that of light)**	**2.36 to** $2.83\times10^{10}\ cm/s$ **(2.36 to** $2.83\times10^8\ m/s$**)**	**Same as that of light** $3\times10^{10}\ cm/s$ $(3\times10^8\ m/s)$
(5) Ionizing power : Very high nearly 100 times to that of β-rays.	**Low nearly 100 times to that of γ-rays.**	**Very low.**

(6) Effect on ZnS plate : They cause luminescence.	**Very little effect.**	**Very little effect.**
(7) Penetrating power : Low	**100 times that of α-particles.**	**10 times that of β-particles.**
(8) Range : Very small (8-12 cm.)	**More that of α-particles.**	**More**
(9) Nature of product : Product obtained by the loss of 1 α-particle has atomic number less by 2 units and mass number less by 4 units.	**Product obtained by the loss of 1 β-particle has atomic number more by 1 unit, without any change in mass number.**	**There is no change in the atomic number as well as in mass number.**

Note : ✹ β-particles originates in the nucleus; they are not orbital electrons.

✹ β-particles having their velocity almost equal to velocity of light are known as hard β-particles and the others having their velocity $\approx 1 \times 10^{10}\ cm\ sec^{-1}$ are called soft ***β*-particles.**

✹ γ-radiation always accompany alpha or beta emissions and thus are emitted after α- and β-decay.

✹ Only one kind of emission at a time is noticed. No radioactive substance emits both α- and β-particles simultaneously.

Theory of radioactive disintegration.

Rutherford and Soddy, in 1903, postulated that radioactivity is a nuclear phenomenon and all the radioactive changes are taking place in the nucleus of the atom. They presented an interpretation of the radioactive processes and the origin of radiations in the form of a theory known as **theory of radioactive disintegration**. The main points of this theory are as follows :

- The atomic nuclei of the radioactive elements are unstable and liable to disintegrate any moment.
- The disintegration is spontaneous, *i.e.*, constantly breaking. The rate of breaking is not affected by external factors like temperature, pressure, chemical combination etc.
- During disintegration, atoms of new elements called daughter elements having different physical and chemical properties than the parent elements come into existence.
- During disintegration, either alpha or beta particles are emitted from the nucleus.

The disintegration process may proceed in one of the following two ways :

(1) ***α*-particle emission :** When an α-particle $({}_2He^4)$ is emitted from the nucleus of an atom of the parent element, the nucleus of the new element, called daughter element possesses atomic mass or atomic mass number less by four units and nuclear charge or atomic number less by 2 units because α-particle has mass of 4 units and nuclear charge of two units.

$$\underset{\text{Atomic mass}: W,\ \text{Atomic number}: Z}{\textbf{Parent element}} \xrightarrow{-\alpha} \underset{W-4,\ Z-2}{\textbf{Daughter element}}$$

Examples are:

$$\underset{\text{(Radium)}}{{}_{88}Ra^{226}} \longrightarrow \underset{\text{(Radon)}}{{}_{86}Rn^{222}} + {}_2He^4 \ ; \quad \underset{\text{(Uranium)}}{{}_{92}U^{238}} \longrightarrow \underset{\text{(Thorium)}}{{}_{90}Th^{234}} + {}_2He^4$$

$$\underset{\text{(Bismuth)}}{{}_{83}Bi^{213}} \longrightarrow \underset{\text{(Thallium)}}{{}_{81}Tl^{209}} + {}_2He^4 \ ; \quad \underset{\text{(Polonium)}}{{}_{84}Po^{215}} \longrightarrow \underset{\text{(Lead)}}{{}_{82}Pb^{211}} + {}_2He^4$$

(2) ***β*-particle emission :** *β*-particle is merely an electron which has negligible mass. Whenever a beta particle is emitted from the nucleus of a radioactive atom, the nucleus of the new element formed possesses the same atomic mass but nuclear charge or atomic number is increased by 1 unit than the parent element. Beta particle emission is due to the result of decay of neutron into proton and electron.

$${}_0n^1 \rightarrow {}_1p^1 + {}_{-1}e^0$$

The electron produced escapes as a beta-particle-leaving proton in the nucleus.

$$\underset{\substack{\text{Atomic mass : } W \\ \text{Atomic number : } Z}}{\textbf{Parent element}} \xrightarrow{-\beta} \underset{\substack{W \\ Z+1}}{\textbf{Daughter element}}$$

Examples are:

$$\underset{\text{(Lead)}}{{}_{82}Pb^{214}} \longrightarrow \underset{\text{(Bismuth)}}{{}_{83}Bi^{214}} + {}_{-1}e^0 ; \quad \underset{\text{(Thorium)}}{{}_{90}Th^{234}} \longrightarrow + \underset{\text{(protoactinium)}}{{}_{91}Pa^{234}} + {}_{-1}e^0 ; \quad {}_{83}Bi^{213} \longrightarrow \underset{\text{(Polonium)}}{{}_{84}Po^{213}} + {}_{-1}e^0$$

Special case : If in a radioactive transformation 1 alpha and 2 beta-particles are emitted, the resulting nucleus possesses the same atomic number but atomic mass is less by 4 units. A radioactive transformation of this type always produces an isotope of the parent element.

$${}_ZA^W \xrightarrow{-\alpha} {}_{Z-2}B^{W-4} \xrightarrow{-\beta} {}_{Z-1}C^{W-4} \xrightarrow{-\beta} {}_ZD^{W-4}$$

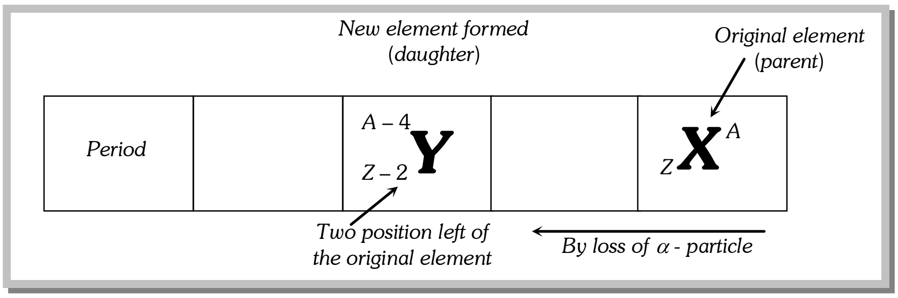
A and *D* are isotopes.

- *γ*-rays are emitted due to secondary effects. The excess of energy is released in the form of *γ*-rays. *Thus γ-rays arise from energy re-arrangements in the nucleus.* As *γ*-rays are short wavelength electromagnetic radiations with no charge and no mass, their emission from a radioactive element does not produce new element.

Group displacement law.

Soddy, Fajans and Russell (1911-1913) observed that when an *α*-particle is lost, a new element with atomic number less by 2 and mass number less by 4 is formed. Similarly, when *β*-particle is lost, new element with atomic number greater by 1 is obtained. The element emitting then *α* or *β*-particle is called **parent** element and the new element formed is called **daughter** element. The above results have been summarized as **Group displacement laws** as follows :

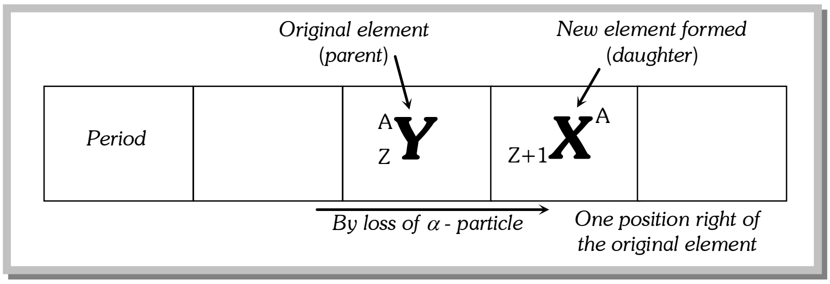
(1) When an *α*-particle is emitted, the new element formed is displaced two positions to the **left** in the periodic table than that of the parent element (*because the atomic number decreases by 2*).



For example, when

$${}_{92}U^{238} \longrightarrow {}_{90}Th^{234} + {}_2He^4$$

(2) When a *β*-particle is emitted, the new element formed is displaced one position to the **right** in the periodic table than that of the parent element (*because atomic number increased by 1*).



For example,

$$_{90}Th^{234} \longrightarrow {}_{91}Pa^{234} + {}_{-1}e^{0};\ {}_{6}C^{14} \longrightarrow {}_{7}N^{14} + {}_{-1}e^{0}$$

Hence, group displacement law should be applied with great care especially in the case of elements of lanthanide series (57 to 71), actinide series (89 to 103), VIII group (26 to 28; 44 to 46; 76 to 78), IA and IIA groups. It is always beneficial to keep in mind the setup and skeleton of the extended form of periodic table.

IA	IIA	IIIB	IVB	VB	VIB	VIIB		VIII		IB	IIB	IIIA	IVA	VA	VIA	VIIA	Zero
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18

	IA 1	IIA	IIIB	IVB	Zero 18
	At. No.				At. No.
1st period	1	–	–	–	2
2nd period	3	–	–	–	10
3rd period	11	–	–	–	18
4th period	19	–	–	–	36
5th period	37	–	–	–	54
6th period	55	56	57*–71	72	86
7th period	87	88	89!–103	104	

*Lanthanides, ! Actinides

Important tips

- *α-Decay produces isodiapher i.e., the parent and daughter nuclide formed by α-decay have same isotopic number, i.e., difference between the number of neutrons and protons is same. For example,*

	$_{88}Ra^{226} \longrightarrow$	$_{86}Rn^{222}$
No. of neutrons	138	136
No. of protons	88	86
Difference	50	50

Thus note that an α-decay leads to

(i) ***Decrease*** *in atomic weight, mass number and number of nucleons by* ***four units***.

(ii) ***Decrease*** *in number of protons, neutrons, nuclear charge and atomic number by* ***two units***.

(iii) ***Increase*** *in n/p ratio.*

- *β-Decay results in the formation of an* ***isobaric element*** *i.e., parent and daughter nuclide have different atomic numbers but same mass number. For example,*

$$_{19}K^{40} \longrightarrow {}_{20}Ca^{40} + {}_{-1}e^0$$

Thus note that a β-decay leads to

(i) ***No change*** *in atomic weights, mass number and number of nucleons.*

(ii) ***Decrease*** *in number of neutrons by* ***one unit****.*

(iii) ***Increase*** *in nuclear charge, number of protons and atomic number by* ***one unit****.*

(iv) ***Decrease*** *in n/p ratio.*

It is important to note that although β-particle (electron) is not present in the nucleus, even then it is emitted from the nucleus since a neutron at first breaks down to a proton and electron.

$$_0n^1 \longrightarrow {}_1p^1 + {}_{-1}e^0$$

The proton is retained by the nucleus while the electron is emitted as a β-particle.

- *Emission of 1 α-particle and 2 β-particles in succession produces an* ***isotope*** *of the parent element. For example,*

$$_{92}U^{235} \xrightarrow{-\alpha} {}_{90}Th^{231} \xrightarrow{-\beta} {}_{91}Pa^{231} \xrightarrow{-\beta} {}_{92}U^{231}$$

- *This law helps to fix the position of the radioelements in the periodic table.*
- *To determine the number of α- and β- particles emitted during the nuclear transformation. It can be done in following manner :*

$$^a_cX \rightarrow {}^b_dY + x\,{}^4_2He + y\,{}_{-1}e^0,$$

$$a = b + 4x \text{ or } \boxed{x = \frac{a-b}{4}} \quad \text{.......(i)}$$

$$c = d + 2x - y \quad \text{......(ii)}$$

where x = no. of α-emitted, y = no. of β-emitted

substituting the value of x from eq. (i) in eq. (ii) we get $c = d + \left(\frac{a-b}{4}\right)2 - y$; $\boxed{y = d + \left[\frac{a-b}{2}\right] - c}$

Example : 4 Calculate the number of neutrons in the remaining atom after emission of an alpha particle from $^{238}_{92}U$ atom **[Roorkee 1978]**

(a) 146 (b) 145 (c) 144 (d) 143

Solution: (c) On account of emission of an alpha particle, the atomic mass is decreased by 4 units and atomic number by 2 units

So, Atomic mass of daughter element = 234

Atomic number of daughter element = 90

Number of neutrons = Atomic mass – Atomic number

= 234 – 90 = 144

Example : 5 Radioactive disintegration of $^{226}_{88}Ra$ takes place in the following manner into RaC

$Ra \xrightarrow{-\alpha} Rn \xrightarrow{-\alpha} RaA \xrightarrow{-\alpha} RaB \xrightarrow{-\beta} RaC$, Determine mass number and atomic number of RaC.

(a) 214 and 84 (b) 214 and 86 (c) 214 and 83 (d) 214 and 85

Solution: (c) Parent element is $^{226}_{88}Ra$

Atomic mass = 226

Atomic number = 88

RaC is formed after the emission of 3 alpha particles. Mass of 3 alpha particles $= 3 \times 4 = 12$

So Atomic mass of $RaC = (226 - 12) = 214$

With emission of one α-particle, atomic number is decreased by 2 and with emission of β-particle, atomic number is increased by 1.

So Atomic number of $RaC = 88 - (3 \times 2) + 1 = 83$

Example : 6 How many 'α and β' particles will be emitted when $_{90}Th^{234}$ changes into $_{84}Po^{218}$ **[CBSE 1992]**

(a) 2 and 6 (b) 4 and 2 (c) 2 and 4 (d) 6 and 2

Solution: (b) The change is; $\underset{\text{Parent}}{_{90}Th^{234}} \rightarrow \underset{\text{End product}}{_{84}Po^{218}}$

Decrease in mass $= (234 - 218) = 16\,amu$

Mass of 1 α-particle $= 4\,amu$

Therefore, number of α-particles emitted $= \frac{16}{4} = 4$

Number of β-particles emitted $= 2 \times$ No. of αparticles emitted $-$ (atomic no. of parent $-$ At. no. of product)

$= 2 \times 4 - (90 - 84) = 2$

Hence number of α-particles = 4 and number of β-particles = 2

Example : 7 An atom has atomic mass 232 and atomic number 90. During the course of disintegration, it emits 2 β-particles and few α-particles. The resultant atom has atomic mass 212 and atomic number 82. How many α-particles are emitted during this process **[CBSE 1992]**

(a) 5 (b) 6 (c) 7 (d) 8

Solution: (a) The decrease in atomic mass $= (232 - 212) = 20$

Decrease in mass occurs due to emission of α-particles. Let x be the number of alpha particles emitted.

Mass of 'x' α-particles $= 4x$

So $4x = 20$ or $x = \frac{20}{4} = 5$

Alternative method : This can also be determined by the application of following equation :

No. of β-particles emitted $= 2 \times$ No. of α-particles emitted $-(Z_{\text{Parent}} - Z_{\text{End product}})$

$2 = 2 \times x - (90 - 82)$ or $x = 5$

Example : 8 An element X with atomic number 90 and mass number 232 loses one α- and two β-particles successively to give a stable species Z. What would be the atomic number and atomic weight of Z **[CPMT 1990]**

(a) $_{90}Z^{228}$ (b) $_{91}Z^{235}$ (c) $_{90}Z^{235}$ (d) $_{91}Z^{238}$

Solution: (a) At. no. and At. wt. of the element (Y) produced by the loss of one α-particle ($_2He^4$) from $_{90}X^{232} = {_{88}Y^{228}}$

At. no. and At.wt. of Z produced by the loss of 2 β-particles ($_{-1}e^0$) from $Y = {_{90}Z^{228}}$

Example : 9 Find out the total number of α- and β-particles emitted in the disintegration of $_{90}Th^{232}$ to $_{82}Pb^{208}$ **[CPMT 1989]**

(a) 6 and 4 (b) 8 and 4 (c) 9 and 6 (d) 2 and 4

Solution: (a) Change in at. wt. $= 232 - 208 = 24\,amu$ (at. wt. unit)

Now since in one α-particle emission, at. wt. is decreased by 4 *amu*, the number of α-emissions for 24 *amu* = 24/4 = 6

Atomic number after 6α-emissions = 90 – 12 = 78 ($\because \alpha = {}_2He^4$)

Increase in atomic number from 78 to the given 82 = 82 – 78 = 4 ($\because \beta$ - particle = ${}_{-1}e^0$)

$\therefore$ No. of β-particle emissions = 4

Example : 10 ${}_{92}U^{235}$ belongs to group III B of the periodic table. It loses one α-particle to form the new element. Predict the position of the new element in the periodic table **[MLNR 1984]**

Solution: Since loss of an α-particle decreases the atomic number of the element by 2, the resulting product should lie two groups to the left of the parent group. However, in the present case the element will remain in the same group of the periodic table because it is an actinide element.

Example : 11 ${}^{234}_{90}Th$ disintegrates to give ${}^{206}_{82}Pb$ as the final product. How many alpha and beta particles are emitted in this process **[IIT 1986]**

(a) 6 (b) 8 (c) 9 (d) 2

Solution: (a)

$$\underset{\text{Parent}}{{}^{234}_{90}Th} \longrightarrow \underset{\text{End product}}{{}^{206}_{82}Pb}$$

Decrease in mass = (234 – 206) = 28

Mass of α-particle = 4

So Number of α-particles emitted = $\frac{28}{4} = 7$

Number of beta particles emitted = 2 × No. of α-particles – (At. no. of parent – At. no. of end product)

= 2 × 7 – (90 – 82) = 6

Rate of radioactive decay.

"*According to the law of radioactive decay, the quantity of a radioelement which disappears in unit time* (*rate of disintegration*) *is directly proportional to the amount present.*"

The law of radioactive decay may also be expressed mathematically.

Suppose the number of atoms of the radioactive element present at the commencement of observation, *i.e.* when $t = 0$ is N_0, and after time *t*, the number of atoms remaining unchanged is N_t, then the rate of decay of atoms is $-\frac{dN_t}{dt}$ (*the word 'd' indicates a very-very small fraction; the negative sign shows that the number of atoms* N_t *decreases as time t increases*)

Now since the change in number of atoms is proportional to the total number of atoms N_t, the relation becomes $\boxed{-\frac{dN_t}{dt} = \lambda N_t}$, where λ is a radioactive constant or decay constant.

- Rate of decay of nuclide is **independent** of temperature, so its energy of activation is zero.
- Since the rate of decay is directly proportional to the amount of the radioactive nuclide present and as the number of undecomposed atom decreases with increase in time, the rate of decay also decreases with the increase in time. Various forms of equation for radioactive decay are,

$$N_t = N_0 e^{-\lambda t}; \ \ \log N_0 - \log N_t = 0.4343\lambda t \ ; \ \log \frac{N_0}{N_t} = \frac{\lambda t}{2.303}$$

$$\boxed{\lambda = \frac{2.303}{t} \log \frac{N_0}{N_t}}$$

where N_0 = Initial number of atoms of the given nuclide, *i.e.* at time 0

N_t = Number of atoms of that nuclide present after time *t.*

λ = Decay constant

Note : ✹ This equation is similar to that of first order reaction, hence we can say that radioactive disintegration are examples of first order reactions. However, unlike first order rate constant (K), the decay constant (λ) is independent of temperature.

Decay constant (λ) : The ratio between the number of atoms disintegrating in unit time to the total number of atoms present at that time is called the decay constant of that nuclide.

Characteristics of decay constant (λ) :

- It is characteristic of a nuclide (not for an element).
- Its units are time^{-1}.
- Its value is always less than one.

Half life and Average life period.

(1) **Half-life period ($T_{1/2}$ or $t_{1/2}$) : Rutherford** in 1904 introduced a constant known as *half-life* period of the radioelement for evaluating its radioactivity or for comparing its radioactivity with the activities of other radioelements. The half-life period of a radioelement is defined, as the time required by a given amount of the element to decay to one-half of its initial value.

Mathematically, $\boxed{t_{1/2} = \frac{0.693}{\lambda}}$

Now since λ is a constant, we can conclude that half-life period of a particular radioelement is independent of the amount of the radioelement. In other words, whatever might be the amount of the radioactive element present at a time, it will always decompose to its half at the end of one half-life period.

Half-life period is a measure of the radioactivity of the element since shorter the half-life period of an element, greater is the number of the disintegrating atoms and hence greater is its radioactivity. The half-life periods or the half-lives of different radioelements vary widely, ranging form a fraction of a second to million of years.

Fraction and Percent of radioactive nuclides left after *n*-Half-Lives

No. of half-lives passed (n)	**Fraction of mass**		**Percent of mass**	
	Decayed	**Left**	**Decayed**	**Left**
0	**0**	**1.0**	**0**	**100**

$\frac{1}{2}$	$\frac{\sqrt{2}-1}{\sqrt{2}}=0.293$	$\frac{1}{\sqrt{2}}=0.707$	**29.3**	**79.7**
1	$\frac{1}{2}=0.50$	$\frac{1}{2}=0.50$	**50**	**50**
2	$\frac{3}{4}=0.75$	$\frac{1}{4}=0.25$	**75**	**25**
3	$\frac{7}{8}=0.875$	$\frac{1}{8}=0.125$	**87.5**	**12.5**
4	$\frac{15}{16}=0.9375$	$\frac{1}{16}=0.0625$	**93.75**	**6.25**
5	$\frac{31}{32}=0.96875$	$\frac{1}{32}=0.03125$	**96.75**	**3.125**
∞	**Total**	**0**	**100**	**0**

Let the initial amount of a radioactive substance be N_0

After one half-life period $(t_{1/2})$ it becomes $=N_0/2$

After two half-life periods $(2t_{1/2})$ it becomes $=N_0/4$

After three half-life periods $(3t_{1/2})$ it becomes $=N_0/8$

After n half life periods $(nt_{1/2})$ it shall becomes $=\left(\frac{1}{2}\right)^n N_0$

Half life periods of some isotopes

Radio isotope	*Half life*	***Radio isotope***	***Half life***
${}^{238}_{92}U$	4.5×10^9 **years**	${}^{32}_{15}P$	**14.3 days**
${}^{230}_{90}Th$	8.3×10^4 **years**	${}^{131}_{53}I$	**8.0 days**
${}^{226}_{88}Ra$	1.58×10^3 **years**	${}^{214}_{84}Po$	1.5×10^{-4} **seconds**
${}^{234}_{90}Th$	**24 days**	${}^{14}_{6}C$	5×10^3 **years**
${}^{59}_{26}Fe$	**44.3 days**	${}^{222}_{86}Rn$	**3.82 days**

Thus, for the total disintegration of a radioactive substance an infinite time will be required.

Time (T)	***Amount of radioactive substance (N)***	***Amount of radioactive substance decomposed*** (N_0-N)
0	(N_0)	**0**
$t_{1/2}$	$\frac{1}{2}N_0=\left(\frac{1}{2}\right)^1 N_0$	$\frac{1}{2}N_0=\left[1-\frac{1}{2}\right]N_0$

$2t_{1/2}$	$\frac{1}{4}N_0 = \left(\frac{1}{2}\right)^2 N_0$	$\frac{3}{4}N_0 = \left[1-\frac{1}{4}\right]N_0$
$3t_{1/2}$	$\frac{1}{8}N_0 = \left(\frac{1}{2}\right)^3 N_0$	$\frac{7}{8}N_0 = \left[1-\frac{1}{8}\right]N_0$
$4t_{1/2}$	$\frac{1}{16}N_0 = \left(\frac{1}{2}\right)^4 N_0$	$\frac{15}{16}N_0 = \left[1-\frac{1}{16}\right]N_0$
$nt/2$	$\left(\frac{1}{2}\right)^n N_0$	$\left[1-\left(\frac{1}{2}\right)^n\right]N_0$

Amount of radioactive substance left after *n* half-life periods

$$N = \left(\frac{1}{2}\right)^n N_0$$

and Total time $T = n \times t_{1/2}$

where *n* is a whole number.

(2) **Average-life period (*T*) :** Since total decay period of any element is infinity, it is meaningless to use the term total decay period (total life period) for radioelements. Thus the term **average life** is used which the following relation determines.

$$\text{Average life } (T) = \frac{\text{Sum of lives of the nuclei}}{\text{Total number of nuclei}}$$

Relation between average life and half-life : Average life (*T*) of an element is the inverse of its decay constant, *i.e.*, $T = \frac{1}{\lambda}$, Substituting the value of λ in the above equation,

$$T = \frac{t_{1/2}}{0.693} = 1.44\, t_{1/2}$$

Thus, **Average life (*T*)** $= 1.44 \times \text{Half life}(T_{1/2}) = \sqrt{2} \times t_{1/2}$

Thus, the average life period of a radioisotope is approximately under-root two times of its half life period.

Note : ✸ This is because greater the value of λ, *i.e.*, faster is the disintegration, the smaller is the average life (*T*).

Example : 12 The isotopes ^{238}U and ^{235}U occur in nature in the ratio of 140 : 1. Assuming that at the time of earth formation, they were present in equal ratio, make an estimation of the age of earth. The half life period of ^{238}U and ^{235}U are 4.5×10^9 and 7.13×10^8 years respectively **[Roorkee 1983]**

(a) 6.04×10^9 years (b) 5.69×10^{10} years (c) 6.69×10^{11} years (d) 6.69×10^9 years

Solution: (a) Let the age of the earth be *t* years

For ^{238}U $\lambda_1 \times t = 2.303 \log \frac{N_0 U_{238}}{N U_{238}}$ ......(i)

For ^{235}U $\lambda_2 \times t = 2.303 \log \frac{N_0 U_{235}}{N U_{235}}$ ......(ii)

Subtracting eq. (ii) from eq. (i)

$$t(\lambda_1 - \lambda_2) = 2.303\left[\log \frac{N_0 U_{238}}{N U_{238}} - \log \frac{N_0 U_{235}}{N U_{235}}\right] = 2.303 \log \frac{N_0 U_{238}}{N U_{235}} \cdot \frac{N U_{235}}{N U_{238}}$$

$$t\left(\frac{0.693}{4.5 \times 10^9} - \frac{0.693}{7.13 \times 10^8}\right) = 2.303 \log \frac{1}{140}$$

$$= -(2.303)(2.1461)$$

$$= -4.9425$$

$$t = 6.04 \times 10^9 \text{ years}$$

Example : 13 Calculate the mass of ^{14}C (half life period = 5720 *years*) atoms which give 3.7×10^7 disintegrations per second **[Roorkee 1987]**

(a) $2.34 \times 10^{-4}\ g$ (b) $2.24 \times 10^{-4}\ g$ (c) $2.64 \times 10^{-4}\ g$ (d) $2.64 \times 10^{-2}\ g$

Solution: (a) Let the mass of ^{14}C atoms be $m\ g$

Number of atoms in $m\ g$ of $^{14}C = \frac{m}{14} \times 6.02 \times 10^{23}$

$$\lambda = \frac{0.693}{\text{Half life}} = \frac{0.693}{5720 \times 365 \times 24 \times 60 \times 60} = 3.84 \times 10^{-12}\ \text{sec}^{-1}$$

We know that $-\frac{dN_t}{dt} = \lambda \cdot N_t$

i.e. Rate of disintegration $= \lambda \times$ No. of atoms

$$3.7 \times 10^7 = \frac{0.693}{5720 \times 365 \times 24 \times 60 \times 60} \times \frac{m}{14} \times 6.02 \times 10^{23} = \frac{3.84 \times 10^{-12} \times m \times 6.02 \times 10^{23}}{14} = 2.24 \times 10^{-4}\ g$$

Example : 14 Prove that time required for 99.9% decay of a radioactive species is almost ten times to its half life period

Solution: We know that $t = \frac{2.303}{\lambda} \log \frac{N_0}{N_t}$

$N_0 = 100$, $N_t = (100 - 99.9) = 0.1$

So, Time required for 99.9% decay $t = \frac{2.303}{\lambda} \log \frac{100}{0.1}$

$$= \frac{2.303}{\lambda} \times 3 \quad ......(i)$$

Half life period $= \frac{0.693}{\lambda}$ ......(ii)

So $\frac{\text{Time requird for 99.9\% decay}}{\text{Half life period}} = \frac{2.303 \times 3}{\lambda} \times \frac{\lambda}{0.693} = 10$

Example : 15 1.0 g of $^{198}_{79}Au\,(t_{1/2} = 65\,hours)$ decay by β-emission to produce mercury **[Roorkee 1984]**

(a) Write the nuclear reaction for the process.

(b) How much mercury will be present after 260 *hours* ?

Solution: (a) $^{198}_{79}Au \rightarrow {}^{198}_{80}Hg + {}^{0}_{-1}e$

(b) No. of half-lives in 260 *hours* $= \frac{260}{65} = 4$

Amount of gold left after 4 half-lives $= \left(\frac{1}{2}\right)^n = \frac{1}{16}$

Amount of gold disintegrated $= 1 - \frac{1}{16} = \frac{15}{16}g$

So Amount of mercury formed $= \frac{15}{16} = 0.9375g$

Example : 16 A certain radio-isotope $^{A}_{Z}X$ (Half life = 10 *days*) decays to $^{A-4}_{Z-2}Y$. If 1 g of atoms of $^{A}_{Z}X$ is kept in sealed vessel, how much helium will accumulate in 20 *days* **[Roorkee 1986]**

(a) 16800 *mL* (b) 17800 *mL* (c) 18800 *mL* (d) 15800 *mL*

Solution: (a) $^{A}_{Z}X \rightarrow {}^{A-4}_{Z-2}Y + {}^{4}_{2}He$

In two half lives, $\frac{3}{4}$ of the isotope $^{A}_{Z}X$ has disintegrated, *i.e.*, $\frac{3}{4}g$ atom of helium has been formed from $\frac{3}{4}g$ atom of $^{A}_{Z}X$

Volume of 1 g atom of helium = 22400 *mL*

Thus, Volume of $\frac{3}{4}g$ atom of helium $= \frac{3}{4} \times 22400\,mL = 16800\,mL$

Example : 17 $^{210}_{84}Po$ decays with α-particle to $^{206}_{82}Pb$ with a half life period of 138.4 days. If 1g of $^{210}_{84}Po$ is placed in a sealed tube, how much helium will be accumulated in 69.2 days ? Express the answer in cm^3 at *STP*. **[Roorkee 1991]**

Solution : $^{210}_{84}Po \rightarrow {}^{206}_{82}Pb + {}^{4}_{2}He$

Amount of polonium left after 69.2 days can be calculated by applying,

$N = No\left(\frac{1}{2}\right)^n$ where $n = \frac{69.2}{138.4} = \frac{1}{2}$

$= 1\left(\frac{1}{2}\right)^{1/2} = 0.7072g$

Amount of polonium disintegrated = (1 – 0.7072) = 0.2928g

No. of atoms of polonium in 0.2928g $= \frac{6.023 \times 10^{23}}{210} \times 0.2928$

Thus, No. of atoms of helium formed $= \dfrac{6.023 \times 10^{23} \times 0.2928}{210}$

Volume of helium collected $= \dfrac{22400}{6.023 \times 10^{23}} \times \text{No. of helium atoms} = \dfrac{22400 \times 0.2928}{210} = 31.23\, cm^3$

Example : 18 In nature a decay chain series starts with ${}^{232}_{90}Th$ and finally terminates at ${}^{208}_{82}Pb$. A thorium ore sample was found to contain $8 \times 10^{-5} mL$ of helium at STP and $5 \times 10^{-7} g$ of ${}^{232}Th$. Find the age of the ore sample assuming the source of helium to be only decay of ${}^{232}Th$. Also assume complete retention of helium within the ore. (Half life of ${}^{232}Th = 1.39 \times 10^{10}$ years) **[Roorkee 1992]**

(a) 6.89×10^9 years (b) 4.89×10^9 years (c) 3.69×10^9 years (d) 6.893×10^{10} years

Solution: (b) No. of moles of helium $= \dfrac{8 \times 10^{-5}}{22400}$

${}^{232}_{90}Th \rightarrow {}^{208}_{82}Pb + 6\, {}^{4}_{2}He$

No. of ${}^{232}_{90}Th$ moles which have disintegrated $= \dfrac{8 \times 10^{-5}}{6 \times 22400}$

Mass of ${}^{232}_{90}Th$ which have disintegrated $= \dfrac{8 \times 10^{-5} \times 232}{6 \times 22400} = 1.3809 \times 10^{-7}\, g$

Mass of ${}^{232}Th$ left, $'N_t' = 5 \times 10^{-7}\, g$

$'N_0' = (5 \times 10^{-7} + 1.3809 \times 10^{-7}) = 6.3809 \times 10^{-7}\, g$

Applying $t = \dfrac{2.303}{\lambda} \log \dfrac{N_0}{N_t} = \dfrac{2.303}{0.693} \times 1.39 \times 10^{10} \log \dfrac{6.3809 \times 10^{-7}}{5 \times 10^{-7}} = 4.89 \times 10^9$ years

Radioactive disintegration series.

The phenomenon of natural radioactivity continues till stable nuclei are formed. All the nuclei from the initial element to the final stable element constitute a series known as disintegration series. Further we know that mass numbers change only when α-particles are emitted (and not when β-particles are emitted) causing the change in mass of 4 units at each step. Hence the mass numbers of all elements in a series will fit into one of the formulae.

$$4n,\ 4n+1,\ 4n+2 \text{ and } 4n+3$$

Hence there can be only four disintegration series

Series	$4n$	$4n+1$	$4n+2$	$4n+3$
n	58	59	**59**	**58**
Parent element	${}_{90}Th^{232}$	${}_{94}Pu^{241}$	${}_{92}U^{238}$	${}_{92}U^{235}$
Half life	1.39×10^{10} **years**	**10 years**	4.5×10^9 **years**	7.07×10^8 **years**
Prominent element	${}_{90}Th^{232}$	${}_{93}Np^{237}$	${}_{92}U^{238}$	${}_{89}Ac^{227}$
Half life	1.39×10^{10} **years**	2.2×10^6 **years**	4.5×10^9 **years**	**13.5 years**
Name of series	**Thorium (Natural)**	**Neptunium (Artificial)**	**Uranium (Natural)**	**Actinium (Natural)**
End product	${}_{82}Pb^{208}$	${}_{83}Bi^{209}$	${}_{82}Pb^{206}$	${}_{82}Pb^{207}$

n	**52**	**52**	**51**	**51**
Number of lost particles	$\alpha = 6$	$\alpha = 8$	$\alpha = 8$	$\alpha = 7$
	$\beta = 4$	$\beta = 5$	$\beta = 6$	$\beta = 4$

The numbers indicate that in a particular series the mass numbers of all the members are either divisible by 4 (in case of 4*n*) or divisible by 4 with remainder of 1, 2 or 3 (in the rest three series), *n* being an integer. In other words, the mass numbers of the members of 4*n*, $4n+1$, $4n+2$ and $4n+3$ series are exactly divisible by 4, 4 + 1, 4 + 2 and 4 + 3 respectively.

Note : ✹ 4*n* + 1 series is an artificial series while the rest three are natural.

✹ The end product in the 4*n* + 1 series is bismuth, while in the rest three, a stable isotope of lead is the end product.

✹ The 4*n* + 1 series starts from plutonium $_{94}Pu^{241}$ but commonly known as neptunium series because neptunium is the longest-lived member of the series.

✹ The 4*n* + 3 series actually starts from $_{92}U^{235}$.

Activity of population, Radioactive equilibrium and Units of radioactivity.

(1) **Activity of population or specific activity :** It is the measure of radioactivity of a radioactive substance. It is defined as ' the number of radioactive nuclei, which decay per second per gram of radioactive isotope.' Mathematically, if '*m*' is the mass of radioactive isotope, then

$$\text{Specific activity} = \frac{\text{Rate of decay}}{m} = \frac{\lambda N}{m} = \lambda \times \frac{\text{Avogadro number}}{\text{Atomic mass in } g}$$

where *N* is the number of radioactive nuclei which undergoes disintegration.

(2) **Radioactive equilibrium :** Suppose a radioactive element *A* disintegrates to form another radioactive element *B* which in turn disintegrates to still another element *C*.

$$A \longrightarrow B \longrightarrow C$$

In the starting, the amount of *A* (in term of atoms) is large while that of *B* is very small. Hence the rate of disintegration of *A* into *B* is high while that of *B* into *C* is low. With the passage of time, *A* go on disintegrating while more and more of *B* is formed. As a result, the rate of disintegration of *A* to *B* goes on decreasing while that of *B* to *C* goes on increasing. Ultimately, a stage is reached when the rate of disintegration of *A* to *B* is equal to that of *B* to *C* with the result the amount of *B* remains constant. Under these conditions *B* is said to be in equilibrium with *A*. For a radioactive equilibrium to be established half-life of the parent must be much more than half-life of the daughter.

It is important to note that the term equilibrium is used for reversible reactions but the radioactive reactions are irreversible, hence it is preferred to say that *B* is in a steady state rather than in equilibrium state.

At a steady state,

$$\boxed{\frac{N_A}{N_B} = \frac{\lambda_B}{\lambda_A} = \frac{T_A}{T_B}\left(\because T = \frac{1}{\lambda}\right)}$$, Where λ_A and λ_B are the radioactive constants for the processes $A \rightarrow B$ and $B \rightarrow C$ respectively. Where T_A and T_B are the average life periods of A and B respectively.

In terms of half-life periods, $\boxed{\frac{N_A}{N_B} = \frac{(t_{1/2})_A}{(t_{1/2})_B}}$

Thus at a steady state (at radioactive equilibrium), the amounts (number of atoms) of the different radioelements present in the reaction series are inversely proportional to their radioactive constants or directly proportional to their half-life and also average life periods.

It is important to note that the radioactive equilibrium differs from ordinary chemical equilibrium because in the former the amounts of the different substances involved are not constant and the changes are not reversible.

(3) **Units of radioactivity :** The standard unit in radioactivity is curie (*c*) which is defined as that amount of any radioactive material which gives 3.7×10^{10} disintegration's per second (*dps*), *i.e.*,

$1c$ = Activity of 1*g* of $Ra^{226} = 3.7 \times 10^{10}\, dps$

The millicurie (*mc*) and microcurie (μc) are equal to 10^{-3} and 10^{-6} curies *i.e.* 3.7×10^7 and $3.7 \times 10^4\, dps$ respectively.

$1c = 10^3\, mc = 10^6\, \mu c$; $1c = 3.7 \times 10^{10}\, dps$; $1mc = 3.7 \times 10^7\, dps$; $1\mu c = 3.7 \times 10^4\, dps$

But now a day, the unit curie is replaced by rutherford (*rd*) which is defined as the amount of a radioactive substance which undergoes $10^6\, dps$. *i.e.*, $1\,rd = 10^6\, dps$. The **millicurie** and **microcurie** correspondingly rutherford units are **millirutherford** (*mrd*) and **microrutherford** (μrd) respectively.

$1\,c = 3.7 \times 10^{10}\, dps = 37 \times 10^3\, rd$; $1\,mc = 3.7 \times 10^7\, dps = 37\, rd$; $1\,\mu c = 3.7 \times 10^4\, dps = 37\, mrd$

However, is SI system the unit of radioactivity is **Becquerel** (*Bq*)

$1\, Bq$ = 1 disintegration per second = 1 *dps* = $1\mu rd$

$10^6\, Bq = 1\, rd$

$3.7 \times 10^{10}\, Bq = 1\, c$

(4) **The Geiger-Nuttal relationship :** It gives the relationship between decay constant of an α- radioactive substance and the range of the α-particle emitted.

$$\boxed{\log \lambda = A + B \log R}$$

Where *R* is the range or the distance which an α-particle travels from source before it ceases to have ionizing power. *A* is a constant which varies from one series to another and *B* is a constant for all series. It is obvious that the greater the value of λ the greater the range of the α-particle.

Artificial transmutation, Synthetic or Transuranic elements and Artificial radioactivity.

(1) **Artificial transmutation or Nuclear transformation or Nuclear transmutation :** The conversion of one element into another by artificial means, *i.e.*, by means of bombarding with some fundamental particles, is

known as artificial transmutation. The phenomenon was first applied on nitrogen whose nucleus was bombarded with α-particles to produce oxygen.

$$\underset{\text{Nitrogen isotope}}{{}_7N^{14}} + \underset{\text{Alpha particle}}{{}_2He^4} \rightarrow \underset{\text{Oxygen isotope}}{{}_8O^{17}} + \underset{\text{Proton}}{{}_1H^1}$$

The element, which is produced, shows radioactivity, the phenomenon is known as **Induced radioactivity**. The fundamental particles which have been used in the bombardment of different elements are as follows :

(i) α-particle : Helium nucleus, represented as ${}_2He^4$.

(ii) Proton : Hydrogen nucleus, represented as ${}_1H^1$.

(iii) Deutron : Deuterium nucleus, represented as ${}_1H^2$ or ${}_1D^2$.

(iv) Neutron : A particle of mass number 1 but no change, represented as ${}_0n^1$.

Since α-particles, protons and deutrons carry positive charge, they are repelled by the positively charged nucleus and hence these are not good projectiles. On the other hand, **neutrons**, which carry no charge at all, are the **best projectiles**. Further among α-particles, protons and deutrons; the latter two carrying single positive charge are better projectiles than the α-particles. However, the positively charged α-particles, protons and deutrons can be made much more effective if they are imparted with high velocity. **Cyclotron** is the most commonly used instrument for accelerating these particles. The particles leave the instrument with a velocity of about 25,000 miles per *second*. A more recent accelerating instrument is called the **synchrotron or bevatron**. It is important to note that this instrument cannot accelerate the neutrons, being neutral.

When a target element is bombarded with neutrons, product depends upon the speed of neutrons. **Slow neutrons** penetrate the nucleus while a **high-speed neutron passes through the nucleus.**

$${}_{92}U^{238} + \underset{\text{slow speed}}{{}_0n^1} \rightarrow {}_{92}U^{239};\ {}_{92}U^{238} + \underset{\text{high speed}}{{}_0n^1} \rightarrow {}_{92}U^{237} + 2{}_0n^1$$

Thus slow neutrons, also called **thermal neutrons** are more effective in producing nuclear reactions than high-speed neutrons.

Alchemy : The process of transforming one element into other is known as alchemy and the person involved in such experiments is called alchemist. Although, gold can be prepared from lead by alchemy, the gold obtained is radioactive and costs very high than natural gold**.**

Some examples are given below for different nuclear reactions :

(i) **Transmutation by α-particles :**

(a) ***α, n type***

$${}_4Be^9(\alpha,n)\,{}_6C^{12} \text{ i.e. } {}_4Be^9 + {}_2He^4 \rightarrow {}_6C^{12} + {}_0n^1;\ {}_{94}Pu^{239}(\alpha,n)\,{}_{96}Cm^{242} \text{ i.e. } {}_{94}Pu^{239} + {}_2He^4 \rightarrow {}_{94}Cm^{242} + {}_0n^1$$

(b) ***α, p type***

$${}_9F^{19}(\alpha,p)\,{}_{10}Ne^{22} \text{ i.e. } {}_9F^{19} + {}_2He^4 \rightarrow {}_{10}Ne^{22} + {}_1H^1;\ {}_7N^{14}(\alpha,p)\,{}_8O^{17} \text{ i.e., } {}_7N^{14} + {}_2He^4 \rightarrow {}_8O^{17} + {}_1H^1$$

(c) ***α, β type***

$_{26}Fe^{59}\,(\alpha,\beta)\,_{29}Cu^{63}$ i.e., $_{26}Fe^{59} + {_2He^4} \rightarrow {_{29}Cu^{63}} + {_{-1}e^0}$

(ii) **Transmutation by protons :**

(a) ***p, n* type**

$_{15}P^{31}\,(p,n)\,_{16}S^{31}$ i.e., $_{15}P^{31} + {_1H^1} \rightarrow {_{16}S^{31}} + {_0n^1}$

(b) ***p, γ* type**

$_6C^{12}\,(p,\gamma)\,_7N^{13}$ i.e., $_6C^{12} + {_1H^1} \rightarrow N^{13} + \gamma$

(c) ***p, d* type**

$_4Be^9\,(p,d)\,_4Be^8$ i.e., $_4Be^9 + {_1H^1} \rightarrow {_4Be^8} + {_1H^2}$

(d) ***p, α* type**

$_8O^{16}\,(p,\alpha)\,_7N^{31}$ i.e., $_8O^{16} + {_1H^1} \rightarrow {_7N^{13}} + {_2He^4}$

(iii) **Transmutation by neutrons :**

(a) ***n,p* type**

$_{13}Al^{27}\,(n,p)\,_{12}Mg^{27}$ i.e., $_{13}Al^{27} + {_0n^1} \rightarrow {_{12}Mg^{27}} + {_1H^1}$

(b) ***n,α* type**

$_8O^{16}\,(n,\alpha)\,_{12}Mg^{27}$ i.e., $_8O^{16} + {_0n^1} \rightarrow {_6C^{13}} + {_2He^4}$

(c) ***n, γ* type**

$_{92}U^{238}\,(n,\lambda)\,_{92}U^{239}$ i.e., $_{92}U^{238} + {_0n^1} \rightarrow {_{92}U^{238}} + \lambda$

(d) ***n,β* type**

$_8O^{18}\,(n,\beta)\,_9F^{19}$ i.e., $_8O^{18} + {_0n^1} \rightarrow {_9F^{19}} + {_{-1}e^0}$

(iv) **Transmutation by deutrons :**

(a) ***d,p* type**

$_3Li^6\,(d,p)\,_3Li^7$ i.e., $_3Li^6 + {_1H^2} \rightarrow {_3Li^7} + {_1H^1}$; $_{32}As^{75}\,(d,p)\,_{32}As^{76}$ i.e., $_{32}As^{75} + {_1H^2} \rightarrow {_{32}As^{76}} + {_1H^1}$

(v) **Transmutation by *γ*-radiations :**

(a) ***γ, n* type**

$_4Be^9\,(\gamma,n)\,_4Be^8$ i.e., $_4Be^9 + \gamma \rightarrow {_4Be^8} + {_0n^1}$

(2) **Synthetic elements :** Elements with atomic number greater than 92 *i.e.* the elements beyond uranium in the periodic table are not found in nature like other elements. All these elements are prepared by artificial transmutation technique and are therefore known as **transuranic elements** or **synthetic elements**. The nuclear reactions for the preparation of some transuranic elements are cited below.

Elements 93 **(neptunium)** and 94 **(plutonium)** were first discovered in 1940. Bombarding uranium-238 with neutrons produced them.

$$\underset{\text{(Uranium)}}{_{92}U^{238}} + {_0n^1} \rightarrow {_{92}U^{239}} \rightarrow \underset{\text{(Neptunium)}}{_{93}Np^{239}} + {_{-1}e^0};\ {_{94}Np^{239}} \rightarrow {_{94}Pu^{239}} \rightarrow {_{-1}e^0}$$

Elements with larger atomic numbers are normally formed in small quantities in particle accelerators. For example, curium-242 is formed when a plutonium-239 target is struck with alpha particles.

$$\underset{\text{Plutonium}}{{}_{94}Pu^{239}} + {}_2He^4 \rightarrow \underset{\text{curium}}{{}_{96}Cm^{242}} + {}_0n^1$$

(3) **Artificial radioactivity or induced radioactivity :** In 1934, **Irene Curie and F. Joliot** observed that when boron and aluminium were bombarded by α-particles, neutrons, protons and positrons were emitted.

Curie and Joliot explained this observation by saying that during bombardment, a metastable isotope is formed which behaves as a radioactive element. This process was termed as **artificial radioactivty**.

"The process in which a stable isotope is converted into radioactive element by artificial transmutation is called artificial radioactivity."

When ${}^{27}_{13}Al$ is bombarded by α-particles, radioactive isotope ${}^{30}_{15}P$ is formed.

$${}^{27}_{13}Al + {}^{4}_{2}He \rightarrow {}^{30}_{14}Si + {}_1H^1 \text{ (95\% of total conversion)}$$

$${}^{27}_{13}Al + {}^{4}_{2}He \rightarrow {}^{30}_{15}P^* + {}_0n^1 \text{ (5\% of total conversion half life period of } {}^{30}_{15}P \text{ is 3.2 minutes)}$$

$${}^{30}_{15}P^* \rightarrow {}^{30}_{14}Si + \underset{\text{Positron}}{{}_{+1}e^0}$$

In a similar manner, the artificial radioactivity was observed when ${}^{10}_{5}B$ was bombarded by α-particles.

$${}^{10}_{5}B + {}^{4}_{2}He \rightarrow {}^{13}_{6}C + {}^{1}_{1}H$$

$${}^{10}_{5}B + {}^{4}_{2}He \rightarrow {}^{13}_{7}N^* + {}_0n^1$$

$${}^{13}_{7}N^* \rightarrow {}^{13}_{6}C + {}_{+1}e^0$$

The following are some of the nuclear reactions in which radioactive isotope are fomed.

${}^{23}_{11}Na + {}^{2}_{1}H \rightarrow {}^{24}_{11}Na^* + {}^{1}_{1}H$ (${}^{24}_{11}Na - \beta$ radioactive); ${}^{238}_{92}U + {}^{1}_{0}n \rightarrow {}^{239}_{92}U^* + \gamma$ (${}^{239}_{92}U - \beta$ radioactive)

${}^{12}_{6}C + {}^{1}_{1}H \rightarrow {}^{13}_{7}N^* + \gamma$ (${}^{13}_{7}N -$ positron radioactive); ${}^{25}_{12}Mg + {}^{4}_{2}He \rightarrow {}^{28}_{13}Al^* + {}^{1}_{1}H$ (${}^{28}_{13}Al - \beta$ radioactive)

Nuclear fission and Nuclear fusion.

(1) **Nuclear fission :** *The splitting of a heavier atom like that of uranium – 235 into a number of fragments of much smaller mass, by suitable bombardment with sub-atomic particles with liberation of huge amount of energy is called **Nuclear fission***. **Hahn and Startsman** discovered that when uranium-235 is bombarded with neutrons, it splits up into two relatively lighter elements.

$${}_{92}U^{235} + {}_0n^1 \rightarrow {}_{56}Ba^{140} + {}_{36}Kr^{93} + 2-3\ {}_0n^1 + \text{ Huge amount of energy}$$

Spallation reactions are similar to nuclear fission. However, they differ by the fact that they are brought by high energy bombarding particles or photons.

Elements capable of undergoing nuclear fission and their fission products. Among elements capable of undergoing nuclear fission, uranium is the most common. The natural uranium consists of three isotopes, namely $U^{234}(0.006\%)$, $U^{235}(0.7\%)$ and $U^{238}(99.3\%)$. Of the three isomers of uranium, nuclear fission of U^{235} and U^{238}

are more important. Uranium-238 undergoes fission by fast moving neutrons while U^{235} undergoes fission by slow moving neutrons; of these two, U^{235} fission is of much significance. Other examples are Pu^{239} and U^{233}.

Uranium-238, the more abundant (99.3%) isotope of uranium, although itself does not undergo nuclear fission, is converted into plutonium-239.

$$_{92}U^{238} + {}_0n^1 \rightarrow {}_{92}U^{239}; \quad {}_{92}U^{239} \rightarrow {}_{93}Np^{239} + {}_{-1}e^0; \quad {}_{93}Np^{238} \rightarrow {}_{94}Pu^{239} + {}_{-1}e^0$$

Which when bombarded with neutrons undergo fission to emit three neutrons per plutonium nucleus. Such material like *U*-238 which themselves are non-fissible but can be converted into fissible material (*Pu*-239) are known as **fertile materials**.

Release of tremendous amount of energy : The importance of nuclear fission lies in the release of tremendous amount of energy during this process. During the U^{235} fission nearly 0.215 mass unit per uranium nucleus is found to be converted into energy.

$$\underbrace{\underset{235.124}{U^{235}} + \underset{1.009}{{}_0n^1}}_{236.133} \rightarrow \underbrace{\underset{138.955}{Xe^{139}} + \underset{94.945}{Sr^{95}} + \underset{2\times1.009}{2\,{}_0n^1} + E}_{235.918}$$

The released energy is due to difference in the total sum of masses of the reactants and products, in according to the Einsten's mass energy relation *i.e.* $E = mc^2$.

Alternatively, $\Delta m = 236.133 - 235.918 = 0.215\,amu$

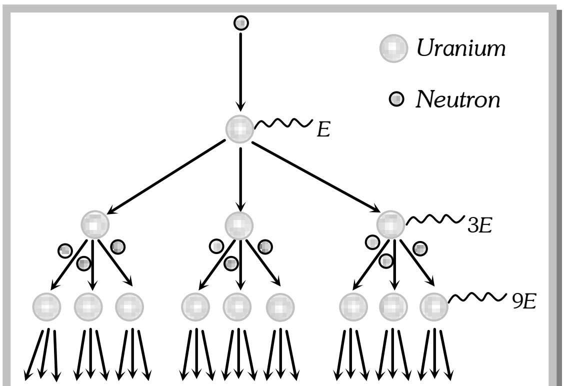
$\because \; 1\,amu = 931\,MeV$

$0.215\,amu = 931 \times 0.215\,MeV = 198\,MeV = 198 \times 2.3 \times 10^7\,kcal$

$\therefore$ Energy released by the fission of 1 *g* of $U^{235} = \dfrac{198 \times 2.3 \times 10^7}{235} = 1.9 \times 10^7\,kcal$

Recall that the combustion of 1 *g* of carbon releases only $94.0/12 = 7.83\,kcal$ of energy while the fission of 1 *g* of U^{235} releases $1.9 \times 10^7\,kcal$. Hence nuclear fission releases several million times higher energy than the ordinary chemical combustion.

Release of neutrons : During U^{235} fission it is obvious that 2-3 neutrons per uranium molecule are emitted. Some neutrons are ejected within an extremely short interval and are called *prompt neutrons*; fission products for an appreciable time fraction of a second to several seconds emit the rest after the fission. These are called *delayed neutrons*.



A fission chain reactor of Uranium.

Note : ✹ *Each fission yields 3 neutrons each of which can cause further fission to give 3 neutrons goes on increasing in geometric progression 1, 3, 9, 27, 81, 243,.... and many geometric progression take place in a very small fraction of a second.*

Chain reaction : With a small lump of U^{235}, most of the neutrons emitted during fission escape but if the amount of U^{235} exceeds a few kilograms (*critical mass*), neutrons emitted during fission are absorbed by adjacent nuclei causing further fission and so producing more neutrons. Now since each fission releases a considerable amount of energy, vast quantities of energy will be released during the chain reaction caused by U^{235} fission.

Atomic bomb : An atomic bomb is based upon the process of that nuclear fission in which no secondary neutron escapes the lump of a fissile material for which the size of the fissile material should not be less than a minimum size called the critical size. There is accordingly a sudden release of a tremendous amount of energy, which represents an explosive force much greater than that of the most powerful TNT bomb. In the world war II in 1945 two atom bombs were used against the Japanese cities of Hiroshima and Nagasaki, the former contained *U*-235 and the latter contained *Pu*-239.

Atomic pile or Nuclear reactor : *It is a device to obtain the nuclear energy in a controlled way to be used for peaceful purposes.* The most common reactor consists of a large assembly of graphite (an allotropic form of carbon) blocks having rods of uranium metal (fuel). Many of the neutrons formed by the fission of nuclei of $_{92}U^{235}$ escape into the graphite, where they are very much slow down (from a speed of about 6000 or more *miles/sec* to a *mile/sec*) and now when these low speed neutrons come back into the uranium metal they are more likely to cause additional fissions. Such a substance likes graphite, which slow down the neutrons without absorbing them is known as a **moderator**. Heavy water, D_2O is another important moderator where the nuclear reactor consists of rods of uranium metal suspended in a big tank of heavy water (swimming pool type reactor). Cadmium or boron are used as control rods for absorbing excess neutrons.

Plutonium from a nuclear reactor : For such purposes the fissile material used in nuclear reactors is the natural uranium which consists mainly (99.3%) of *U*-238. In a nuclear reactor some of the neutrons produced in *U*-235 (present in natural uranium) fission converts *U*-238 to a long-lived plutonium isotope, *Pu*-239 (another fissionable material). Plutonium is an important nuclear fuel. Such reactors in which neutrons produced from fission are partly used to carry out further fission and partly used to produce some other fissionable material are called **Breeder reactors**.

Production of radioactive isotopes by bombarding with neutrons from a nuclear reactor : These radioactive isotopes are used in medicine, industry and hospitals.

Nuclear reactors in India : India is equipped with the five nuclear reactors, namely Apsara (1952), Cirus (1960), Zerlina (1961), Purnima (1972) and *R*-5. Purnima uses plutonium fuel while the others utilize uranium as fuel.

Apsara the first nuclear reactor was completed on 14th August 1952 at Trombay under the guidance of the late Dr. H.J. Bhabha. It is the swimming pool reactor, which consists of a lattice of enriched uranium (fuel) immersed in a large pool of water. Water acts as a moderator, coolant and shield. This reactor is simple, safe, flexible, easily accessible and cheap.

(2) **Nuclear fusion :** *" Oposite to nuclear fission, nuclear fusion is defined as a process in which lighter nuclei fuse together to form a heavier nuclei*. However, such processes can take place at reasonable rates only at very high temperatures of the order of several million degrees, which exist only in the interior of stars. Such processes are, therefore, called **Thermonuclear reactions** (temperature dependent reactions). Once a fusion reaction initiates, the energy released in the process is sufficient to maintain the temperature and to keep the process going on.

$$\underset{\text{Hydrogen}}{4\,_1H^1} \rightarrow \underset{\text{Helium}}{_2He^4} + \underset{\text{Positron}}{2\,_{+1}e^0} + \text{Energy}$$

This is not a simple reaction but involves a set of the thermonuclear reactions, which take place in stars including sun. In other words, *energy of sun is derived due to nuclear fission*.

Calculation of energy released in nuclear fusion : Let us write the reaction involving the fusion of four hydrogen nuclei to form helium nucleus.

$$\underset{\substack{\text{Hydrogen} \\ \text{Mass}: 4\times 1.008144 \\ \text{or } =4.032576}}{4\,{}_1H^1} \rightarrow \underset{\substack{\text{Helium} \\ 4.003873}}{{}_2He^4} + \underset{\substack{\text{Positron} \\ 2\times 0.000558 = 0.001116}}{2\,{}_{+1}e^0} \quad (\text{total } 4.004989)$$

$\therefore$ Loss is mass, $\Delta m = 4.032576 - 4.004989 = 0.027587\,amu$

$\therefore$ Energy released $= 0.027587 \times 931\,MeV = 26.7\,MeV$

$\therefore$ Energy released/*gm* of hydrogen consumed $= \dfrac{26.7}{4} = 6.7\,MeV = 6.7 \times 2.3 \times 10^7\,kcal = 1.54 \times 10^8\,kcal$

Controlled nuclear fusion : Unlike the fission process, the fusion process could not be controlled. Since there are estimated to be some 10^{17} pounds of deuterium (${}_1H^2$) in the water of the earth, and since each pound is equivalent in energy to 2500 tonnes of coal, a controlled fusion reactor would provide a virtually inexhaustible supply of energy.

Comparision of nuclear fission and nuclear fusion : Now let us compare the efficiency of the energy conversion of the two processes, *i.e.* nuclear fission and nuclear fusion

Nuclear fission reaction, ${}_{92}U^{235} + {}_0n^1 \rightarrow {}_{56}Ba^{141} + {}_{36}Kr^{92} + 2-3\,{}_0n^1 + 200\,MeV$

If one atom of uranium is fissioned by one neutron, the percent efficiency in terms of mass converted into energy (where 1 mass unit = 931 *MeV*) will be : $\dfrac{200\,MeV}{(235+1)\text{ mass units} \times 931} \times 100 = 0.09\%$

Nuclear fusion reaction, ${}_1H^2 + {}_1H^3 \rightarrow {}_2He^4 + {}_0n^1 + 17.8\,MeV$

The percent efficiency of the reaction $= \dfrac{17.8\,MeV}{(2+3\text{ mass units}) \times 931} \times 100 = 0.35\%$

Thus it indicates that for these two fission and fusion reactions the percent efficiency is approximately four times greater for the fusion reaction.

Hydrogen bomb : Hydrogen bomb is based on the *fusion of hydrogen nuclei into heavier ones by the thermonuclear reactions with release of enormous energy*.

As mentioned earlier the above nuclear reactions can take place only at very high temperatures. Therefore, it is necessary to have an external source of energy to provide the required high temperature. For this purpose, the atom bomb, (*i.e.*, *fission bomb*) is used as a *primer,* which by exploding provides the high temperature necessary for successful working of hydrogen bomb (*i.e.*, *fusion bomb*). In the preparation of a hydrogen bomb, a suitable quantity of deuterium or tritium or a mixture of both is enclosed in a space surrounding an ordinary atomic bomb. The first hydrogen bomb was exploded in November 1952 in Marshall Islands; in 1953 Russia exploded a powerful hydrogen bomb having power of 1 million tonnes of TNT

A hydrogen bomb is far more powerful than an atom bomb. Thus if it were possible to have sufficiently high temperatures required for nuclear fusion, the deuterium present in sea (as D_2O) sufficient to provide all energy requirements of the world for millions of years.

Note : ✸ The first nuclear reactor was assembled by Fermi in 1942.

Difference between Nuclear fission and fusion

Nuclear fission	*Nuclear fusion*
The process occurs only in the nuclei of heavy elements.	**The process occurs only in the nuclei of light elements.**
The process involves the fission of the heavy nucleus to the lighter nuclei of comparable masses.	**The process involves the fission of the lighter nuclei to heavy nucleus.**
The process can take place at ordinary temperature.	**The process takes place at higher temperature** $10^8\ ^oC$) **.**
The energy liberated during this process is high (200 MeV per fission)	**The energy liberated during the process is comparatively low (3 to 24 MeV per fusion)**
Percentage efficiency of the energy conversion is comparatively less.	**Percentage efficiency of the energy conversion is high (four times to that of the fission process).**
The process can be controlled for useful purposes.	**The process cannot be controlled.**

Isotopes, Isobars, Isotones, Isodiaphers, Isoelectronic species, Isosters and Nuclear isomers.

(1) **Isotopes :** Atoms of a given element which have same atomic number (*nuclear charge*) but *different mass number are called isotopes*. In other words, isotopes are the atoms of the same element differing in mass number. Thus isotopes have same number of protons and electrons but different number of neutrons. They have same position in the periodic table, same chemical properties and same atomic charge. The term was first coined by **Soddy**. However, Aston using mass spectrometer first separated isotopes (Ne^{20} and Ne^{22}).

Examples : (i) $\underset{\substack{\text{Hydrogen (Protium)} \\ (p=1,e=1,n=0)}}{{}_1H^1}$, $\underset{\substack{\text{Deuterium} \\ (p=1,e=1,n=0)}}{{}_1H^2}$, $\underset{\substack{\text{Tritium} \\ (p=1,e=1,n=0)}}{{}_1H^3}$ (ii) ${}_6C^{12}, {}_6C^{13}$ and ${}_6C^{14}$

(iii) ${}_8O^{16}, {}_8O^{17}, {}_8O^{18}$ (iv) ${}_{17}Cl^{35}$ and ${}_{17}Cl^{37}$

Of all the elements, tin has maximum number of stable isotopes (ten).

The fractional atomic weight (35.5) of chlorine is due to the fact that in the ordinary chlorine atom, Cl^{35} and Cl^{37} are present in the ratio of 3 : 1.

$$\therefore \text{ Average atomic weight of } Cl = \frac{3 \times 35 + 1 \times 37}{4} = 35.5\, amu$$

The percentage of a given isotope in the naturally occurring sample of an element is called **Isotopic abundance**. As the isotopic abundance of an element is constant irrespective of its source, atomic weight of an element is constant.

(2) **Isobars :** *Isobars are the atoms of different elements with the same mass number but different atomic numbers*. In other words, isobars have different number of protons, neutrons and electrons but the sum of protons and neutrons (*i.e.*, number of nucleons) is same.

Examples : (i) ${}_{18}Ar^{40}, {}_{19}K^{40}$ and ${}_{20}Ca^{40}$ (ii) ${}_{52}Te^{130}, {}_{54}Xe^{130}$ and ${}_{56}Ba^{130}$.

Since isobars are the atoms of different elements, they will have different physical and chemical properties.

(3) **Isotones :** *Isotones are the atoms of different elements with the same number of neutrons but different mass numbers, e.g.* $_{14}Si^{30}$, $_{15}P^{31}$ and $_{16}S^{32}$. Since the variable factor in isotones is the number of protons (atomic number), they must have different physical and chemical properties.

Examples : (i) $_{14}Si^{30}$, $_{14}P^{31}$ and $_{16}S^{32}$ (ii) $_{19}K^{39}$ and $_{20}Ca^{40}$

(iii) $_{1}H^{3}$ and $_{2}He^{4}$ (iv) $_{6}C^{13}$ and $_{7}N^{14}$

(4) **Isodiaphers :** Atoms having same **isotopic number** are called **isodiaphers**.

Mathematically, isotopic number (isotopic excess) = $(N - Z)$ or $(A - 2Z)$

Where, N = Number of neutrons; Z = Number of protons

Examples : (i) $_{92}U^{235}$ and $_{90}Th^{231}$ (ii) $_{19}K^{39}$ and $_{9}F^{19}$ (iii) $_{29}Cu^{65}$ and $_{24}Cr^{55}$

(5) **Isoelectronic species :** Species (atoms, molecules or ions) having same number of electrons are called **isoelectronic.**

Examples : (i) $N^{3-}, O^{2-}, F^{-}, Ne, Na^{+}, Mg^{2+}, Al^{3+}, CH_4, NH_3, H_2O$ and HF have 10 electrons each.

(ii) $P^{3-}, S^{2-}, Cl^{-}, Ar, K^{+}$ and Ca^{2+} have 18 electrons each.

(iii) H^{-}, He, Li^{+} and Be^{2+} have 2 electrons each.

(iv) CO, CN^{-} and N_2 have 14 electrons each.

(v) N_2O, CO_2 and CNO^{-} have 22 electrons each.

(6) **Isosters :** *Molecules having same number of atoms and also same number of electrons are called isosters.*

Examples : (i) N_2 and CO (ii) CO_2 and N_2O (iii) HCl and F_2

(iv) CaO nad MgS (v) C_6H_6 (benzene) and inorganic benzene B_6N_6.

(7) **Nuclear isomers :** *Nuclear isomers (isomeric nuclei) are the atoms with the same atomic number and same mass number but with different radioactive properties.* They have same number of electrons, protons and neutrons. An example of nuclear isomers is uranium-X (half-life 1.4 min) and uranium-Z (half-life 6.7 hours). **Otto Hahn** discovered nuclear isomers.

The reason for nuclear isomerism is the different energy states of the two isomeric nuclei. One may be in the ground state whereas the other should be in an excited state. The nucleus in the excited state will evidently have a different half-life.

Now-a-days as much as more than 70 pairs of nuclear isomers have been found. Few examples areas follows

(i) $\underset{(T_{1/2}=13.8\,hour)}{^{69}Zn}$ and $\underset{(T_{1/2}=57\,\text{min})}{^{69}Zn}$ (ii) $\underset{(T_{1/2}=4.4\,hour)}{^{80}Br}$ and $\underset{(T_{1/2}=18\,\text{min})}{^{80}Br}$

Example : 19 Naturally occurring boron consists of two isotopes whose atomic weights are 10.01 and 11.01. The atomic weight of natural boron is 10.81. Calculate the percentage of each isotope in natural boron

[IIT 1978, 82; MLNR 1994]

(a) 80 (b) 90 (c) 70 (d) 20

Solution: (a) Let the % of isotope with at. wt. 10.01 = x

$\therefore$ % of isotope with at. wt. 11.01 = $(100 - x)$

Now since, $\text{At. wt.} = \dfrac{x \times 10.01 + (100 - x) \times 11.01}{100}$

$$10.81 = \frac{x \times 10.01 + (100 - x) \times 11.01}{100}$$

$x = 20$

Hence, % of isotope with at. wt. 10.01 = 20

$\therefore$ % of isotope with at. wt. 11.01 = 100 – 20 = 80

Application of radioactivity and Hazards of radiations.

Radioisotopes find numerous applications in a variety of areas such as medicine, agriculture, biology, chemistry, archeology, engineering and industry. Some of the are given below :

(1) **Age determination (carbon dating) :** Radioactive decay follows a very exact law, and is virtually unaffected by heat, pressure, or the state of chemical combination of the decaying nuclei, it can be used as a very precise clock for dating past events. For instance, the age of earth has been determined by uranium dating technique as follows. Samples of uranium ores are found to contain Pb^{206} as a result of long series of α- and β-decays. Now if it is assumed that the ore sample contained no lead at the moment of its formation, and if none of the lead formed from U^{238} decay has been lost then the measurement of the Pb^{206}/U^{238} ratio will give the value of time t of the mineral.

$$\frac{\text{No. of atoms of } Pb^{206}}{\text{No. of atoms of } U^{238} \text{ left}} = e^{-\lambda t - 1}, \quad \text{where } \lambda \text{ is the decay constant of uranium-238}$$

Alternatively, $$\boxed{t = \frac{2.303}{\lambda} \log \frac{\text{Initial amount of } U^{238}}{\text{Amount of } U^{238} \text{ in the mineral present till date}}}$$

Similarly, the less abundant isotope of uranium, U^{235} eventually decays to $Pb^{207}; Th^{232}$ decays to Pb^{208} and thus the ratios of Pb^{207}/U^{235} and Pb^{208}/Th^{232} can be used to determine the age of rocks and minerals. Many ages have been determined this way to give result from hundreds to thousands of million of years,.

Besides the above long-lived radioactive substances *viz.* U^{238}, U^{235} and Th^{232} (which have been present on earth since the elements were formed), several short-lived radioactive species have been used to determine the age of wood or animal fossils. One of the most interesting substances is ${}_6C^{14}$ (half-life 5760 years) which was used by **Willard Libby** (Nobel lauret) in determining the age of carbon-bearing materials (*e.g.* wood, animal fossils, etc.) Carbon-14 is produced by the bombardment of nitrogen atoms present in the upper atmosphere with neutrons (from cosmic rays).

$$_7N^{14} + {_0n^1} \rightarrow {_6C^{14}} + {_1H^1}$$

Thus carbon-14 is oxidised to CO_2 and eventually ingested by plants and animals. The death of plants or animals puts an end to the intake of C^{14} from the atmosphere. After this the amount of C^{14} in the dead tissues starts decreasing due to its disintegration.

$$_6C^{14} \rightarrow {_7N^{14}} + {_{-1}e^0}$$

It has been observed that on an average, one gram of radioactive carbon emits about 12 β-particles per minute. Thus by knowing either the amount of C-14 or the number of β-particles emitted per minute per gram of carbon at the initial and final (present) stages, the age of carbon material can be determined by using the following formulae.

$$\boxed{\lambda = \frac{2.303}{t}\log\frac{N_0}{N_t}} \text{ or } \boxed{t = \frac{2.303}{\lambda}\log\frac{N_0}{N_t}}$$

where t = Age of the fossil, λ = Decay constant, N_0 = Initial radioactivity (in the fresh wood), N_t = Radioactivity in the fossil

The above formula can be modified as

$$t = \frac{2.303}{\lambda}\log\frac{\text{Initial ratio of } C^{14}/C^{12} \text{ (in fresh wood)}}{C^{14}/C^{12} \text{ ratio in the old wood}}$$

$$= \frac{2.303}{\lambda}\log\frac{\text{Initial amount of } C^{14}/C^{12} \text{ (in fresh wood)}}{\text{Amount of } C^{14} \text{ in the old wood}} = \frac{2.303}{\lambda}\log\frac{\text{Radioactivity in fresh wood due to } C^{14}}{\text{Radioactivity in old wood due to } C^{14}}$$

$$= \frac{2.303 \times T_{1/2} \text{ of } C^{14}}{0.693}\log\frac{\text{counts min}^{-1}\ g^{-1} \text{ of } C^{14} \text{ in fresh wood}}{\text{counts min}^{-1}\ g^{-1} \text{ of } C^{14} \text{ in old wood}}$$

Similarly, tritium $_1H^3$ has been used for dating purposes.

(2) **Radioactive tracers (use of radio–isotopes) :** A radioactive isotope can be easily identified by its radioactivity. Because of similar physical and chemical properties of radioisotopes and non-radioisotopes of an element, if a small quantity of the former is mixed with normal isotope, then chemical reactions can be studied by determining the radioactivity of the radioisotope. The radioactivity can, therefore act as a tag or label that allows studying the behaviour of the element or compounding which contains this isotope. An isotope added for this purpose is known as isotopic tracer. The radioactive tracer is also known as an isotopic tracer. The radioactive tracer is also known as an indicator because it indicates the reaction. Radioisotopes of moderate half-life periods are used for tracer work. The activity of radioisotopes can be detected by means of electroscope, the electrometer or the Geiger-Muller counter. Tracers have been used in the following fields :

(i) **In medicine :** Radioisotopes are used to diagnose many diseases. For example, Arsenic – 74 tracer is used to detect the presence of tumours, Sodium –24 tracer is used to detect the presence of blood clots and Iodine –131 tracer is used to study the activity of the thyroid gland. It should be noted that the radioactive isotopes used in medicine have very short half-life periods.

(ii) **In agriculture :** The use of radioactive phosphorus ^{32}P in fertilizers has revealed how phosphorus is absorbed by plants. This study has led to an improvement in the preparation of fertilizers. ^{14}C is used to study the kinetics of photo synthesis.

(iii) **In industry :** Radioisotopes are used in industry to detect the leakage in underground oil pipelines, gas pipelines and water pipes. Radioactive isotopes are used to measure the thickness of materials, to test the wear and tear inside a car engine and the effectiveness of various lubricants. Radioactive carbon has been used as a tracer in studying mechanisms involved in many reactions of industrial importance such as alkylation, polymerization, catalytic synthesis etc.

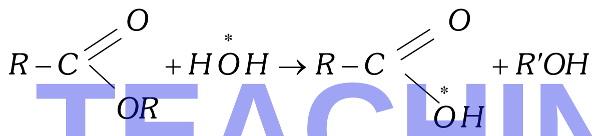
(iv) **Analytical Studies :** Several analytical procedures can be used employing radioisotopes as tracers.

(a) **Adsorption and occlusion studies :** A small amount of radioactive isotope is mixed with the inactive substance and the activity is studied before and after adsorption. Fall in activity gives the amount of substance adsorbed.

(b) **Solubility of sparingly soluble salt :** The solubility of lead sulphate in water may be estimated by mixing a known amount of radioactive lead with ordinary lead. This is dissolved in nitric acid and precipitate as lead sulphate by adding sulphuric acid. Insoluble lead sulphate is filtered and the activity of the water is measured. From this, the amount of $PbSO_4$ still present in water can be estimated.

(c) **Ion-exchange technique :** Ion exchange process of separation is readily followed by measuring activity of successive fractions eluted from the column.

(d) **Reaction mechanism :** By labelling oxygen of the water, mechanism of ester hydrolysis has been studied.



(e) **Study of efficiency of analytical separations :** The efficiency of analytical procedures may be measured by adding a known amount of radio-isotopes to the sample before analysis begins. After the completion, the activity is again determined. The comparison of activity tells about the efficiency of separation.

(3) **Use of $\gamma-$rays :** $\gamma-$rays are used for disinfecting food grains and for preserving food stuffs. Onions, potatoes, fruits and fish etc., when irradiated with $\gamma-$rays, can be preserved for long periods. High yielding disease resistant varieties of wheat, rice, groundnut, jute etc., can be developed by the application of nuclear radiations. The $\gamma-$rays radiations are used in the treatment of cancer. The $\gamma-$ radiations emitted by cobalt –60 can burn cancerous cells. The $\gamma-$ radiations are used to sterilize medical instruments like syringes, blood transfusion sets. etc. These radiations make the rubber and plastics objects heat resistant.

Hazards of radiations : The increased pace of synthesis and use of radio isotopes has led to increased concern about the effect of radiations on matter, particularly in biological systems. Although the radioisotopes have found wide spread uses to mankind such as atomic power generation, dating, tracer technique, medicinal treatment, the use of nuclear energy is an extremely controversial social and political issue. You should ask yourself, how you would feel about having a nuclear power plant in your town. The accident of Chernobyl occurred in 1986 in USSR is no older when radioisotopes caused a hazard there. The nuclear radiations (alpha, beta, gamma as well as X-rays) possess energies far in excess of ordinary bond energies and ionisation energies. Consequently, these radiations are able to break up and ionise the molecules present in living organisms if they are exposed to such radiations. This disrupts the normal functions of living organisms. The damage caused by the radiations, however, depends upon the radiations received. We, therefore, conclude this chapter by examining the health hazards associated with radioisotopes.

The resultant radiation damage to living system can be classified as :

(i) **Somatic or pathological damage :** This affects the organism during its own life time. It is a permanent damage to living civilization produced in body. Larger dose of radiations cause immediate death whereas smaller doses can cause the development of many diseases such as paralysis, cancer, leukaemia, burns, fatigue, nausea, diarrhoea, gastrointestinal problems etc. some of these diseases are fatal.

Many scientists presently believe that the effect of radiations is proportional to exposure, even down to low exposures. This means that any amount of radiation causes some finite risk to living civilization.

(ii) **Genetic damage :** As the term implies, radiations may develop genetic effect. This type of damage is developed when radiations affect genes and chromosomes, the body's reproductive material. Genetic effects are more difficult to study than somatic ones because they may not become apparent for several generations.