

# Introduction to Nuclear Reactor Physics



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# References

- The recommended source of fundamental information about nuclear physics applicable to nuclear reactors are DOE Fundamentals Handbooks:
  - Nuclear Physics and Reactor Theory, Vol. 1 and 2
  - These references are freely available from the DOE website
- There are numerous books dealing with fundamentals of reactor and neutron physics, recommended are:
  - John R. Lamarsh – Introduction to Nuclear engineering
  - Paul Reuss – Neutron Physics
- Multimedia CD available from IAEA
  - Multimedia on Nuclear Reactor Physics



# Elementary Particles

- There are more than 100 elementary particles, most of them with anti-particles
- The name *elementary particles* has more historical meaning
- Many particles believed to be elementary were found to be composed of *quarks*
- Particles without inner structure:
  - fermions – quarks, leptons (electron)
  - bosons
- Composite particles
  - hadrons – baryons (proton, neutron), mezon
- Reactor physics operates mostly with subatomic particles



# Atom

- Atom is the smallest amount of matter that retains the properties of an element
- Atoms are composed of smaller particles that no longer have the properties of the overall element
- One element of the other is distinguished by the kind of atom of which it consists
- Since the atom nature proof in 1803 it was believed that atoms are the elementary particles
- Discovery of radiation indicated that there are subatomic particles



# Subatomic Particles

**Proton** was discovered in 1886

- proton is a positively charged particle with mass  $1.673 \times 10^{-27}$  kg
- charge of a proton is the same in magnitude but opposite in sign to an electron

**Electron** was discovered in 1897

- electrons are negatively charged particles
- mass of an electron is  $9.110 \times 10^{-31}$  kg

**Neutron** was discovered in 1932

- neutron is a neutral particle with mass  $1.675 \times 10^{-27}$  kg
- free neutron is unstable with half-life about 10.4 min.

# Elements and Isotopes

- Atoms of a particular element must have the same number of protons
- Isotopes are variants of an atom with different number of neutrons
- The following notation is used:



Z – Atomic number, number of protons

N – Neutron number, number of neutrons

A – Mass number,  $A = N + Z$ , number of nucleons

					14	15	16	17	18		
					F	F	F	F	F		
				12	13	14	15	16	17		
				O	O	O	O	O	O		
			10	11	12	13	14	15	16		
			N	N	N	N	N	N	N		
			9	10	11	12	13	14	15		
			C	C	C	C	C	C	C		
			7	8	9	10	11	12	13	14	
			B	B	B	B	B	B	B	B	
			6	7	8	9	10	11	12	13	
			Be	Be	Be	Be	Be	Be	Be	Be	
			4	5	6	7	8	9	10	11	12
			Li	Li	Li	Li	Li	Li	Li	Li	Li
			3	4	5	6	7	8	9	10	
			He	He	He	He	He	He	He	He	
			1	2	3	4	5	6			
			H	H	H	H	H	H			
			n								



## Example for Carbon

- Carbon-12, carbon-13 and carbon-14 are three isotopes of the element carbon
- All of them have 6 protons in the nucleus and 6 electrons in orbitals
- They differ in a mass number – 12, 13, and 14, respectively
- There are 6, 7 or 8 neutrons in the nucleus
- These isotopes are designated as  $^{12}\text{C}$ ,  $^{13}\text{C}$ , and  $^{14}\text{C}$





# Atom and Atom Nucleus Dimensions

- Dimensions of electron orbitals are not fixed
- Common definition gives an average distance between the last electron and the nucleus. Typical value is (except for the lightest elements)  $2 \times 10^{-10}$  m
- Density of the electron orbitals is increasing for heavier elements
- Also the dimensions of an atom nucleus are not fixed
- It was determined by measurement of neutrons scattered by a nucleus that radius can be very roughly approximated by the following equation (not valid for light elements):

$$R = 1.25 A^{1/3} \times 10^{-15} \text{ m , where } A \text{ is atom mass number}$$

- Density of a nucleus remains the same with increasing mass number – fundamental requirement for *liquid drop model of atom nucleus*



# Basic Units and Constants

- Nuclear physics required for nuclear reactors operates mostly with atoms and neutrons
- Energy of individual reactions is small, therefore it is convenient to use a special energy unit  $eV$

$$1 \text{ eV} = 1.602 \times 10^{-19} \text{ J}$$

- Mass on atomic scale is expressed by *unified atomic mass unit* –  $u$ 
  - it is by definition 1/12 of the rest mass of an unbound atom of  $^{12}\text{C}$
- Avogadro number expresses the number of elementary entities per 1 mole of a substance
  - equals to number of atoms in 0.012 kg of  $^{12}\text{C}$

$$N_A = 6.022 \times 10^{23} \text{ mol}^{-1}$$



# Atomic Mass and Atomic Weight

**Atomic Mass** is mass of a specific isotope. Can be expressed relatively in unified atomic mass units ( $m_a$ )

**Atomic Weight** is average mass of atoms of an element expressed in  $u$  ( $A_r$ )

**Molar Mass** is mass of one mole of a substance ( $M$ )

- These quantities are often confused
- Relative atomic weight is given in periodic table of elements
- Molar mass is calculated from atomic weight as:

$$M = A_r \times 1 \text{ g/mol}$$



# Quantities Describing Material Properties

- Properties of a material depend on its composition
- Individual components of the material can be expressed in mass fractions ( $w_i$ ) or atomic fractions ( $a_i$ )

$$w_i = \frac{m_i}{m}$$
$$a_i = \frac{N_i}{N}$$

- Number of atoms in 1 cm<sup>3</sup> of matter is expressed by *number density*

$$N = \frac{\rho N_A}{M}$$



# Isotopic Mixtures

- Elements in the nature are occurring in isotopic mixtures
  - For example natural oxygen is a mixture of three isotopes,  $^{16}\text{O}$ ,  $^{17}\text{O}$  and  $^{18}\text{O}$ , atom percent abundance of  $^{16}\text{O}$  is 99.76 %, for  $^{17}\text{O}$  it is 0.04 % and 0.20 % for  $^{18}\text{O}$
- Molar mass (atomic weight) is weighted mean of the molar masses (atomic weights) weighted by isotopic abundance

$$M = \sum_i a_i M_i \quad (1-1)$$

- Example for lithium
  - lithium has isotopes  $^6\text{Li}$  with atom percent abundance 7.5 % and atomic mass 6.015122 u and  $^7\text{Li}$ , 92.5 %, 7.016003 u
  - atomic mass of natural lithium ( $M_{\text{Li}}$ ) can be calculated as:  
 $0.075 \times 6.015122 \text{ u} + 0.925 \times 7.016003 \text{ u} = 6.9409 \text{ u}$



# Conversion of Mass and Atomic Fractions

- Composition of mixtures is usually given in mass or atomic fractions
- Atomic fraction can be calculated from mass fraction as:

$$a_i = w_i \times \frac{M}{M_i} \quad (1-2)$$

- Given the above equation, molar mass of an isotopic mixture can be calculated using mass fractions as:

$$\frac{1}{M} = \sum_i \frac{w_i}{M_i} \quad (1-3)$$



## Mass and Energy

- Equivalence between mass and energy is one of results of Einstein theory of relativity
- Rest energy of a particle corresponds to its rest mass

$$E_0 = m_0 c^2$$

- From the point of view of the theory of relativity, total energy is a sum of rest energy and kinetic energy

$$E = mc^2 = E_0 + E_{kin} = m_0 c^2 + E_{kin}$$

- Relativistic effect are important for particles velocities higher than  $0.2c$ :

$$E \geq 0.02E_0$$

- Given the typical neutron energies in nuclear reactors it is usually not necessary to take into account the relativistic effects for neutrons



# Equivalence of Mass and Energy

- On the nuclear level it is possible to convert between mass and energy
- Sum of energy and mass must be conserved
- Careful measurements have shown that the mass of a particular atom is always slightly less than the sum of the masses of the individual neutrons, protons, and electrons of which the atom consists
- Difference between the mass of the atom and the sum of the masses of its parts is called *mass defect* –  $\Delta m$

$$\Delta m = [Z(m_p + m_e) + (A - Z)m_n] - m_{\text{atom}} \quad (1-4)$$

where  $m_p$ ,  $m_e$  and  $m_n$  are masses of proton, electron and neutron, respectively

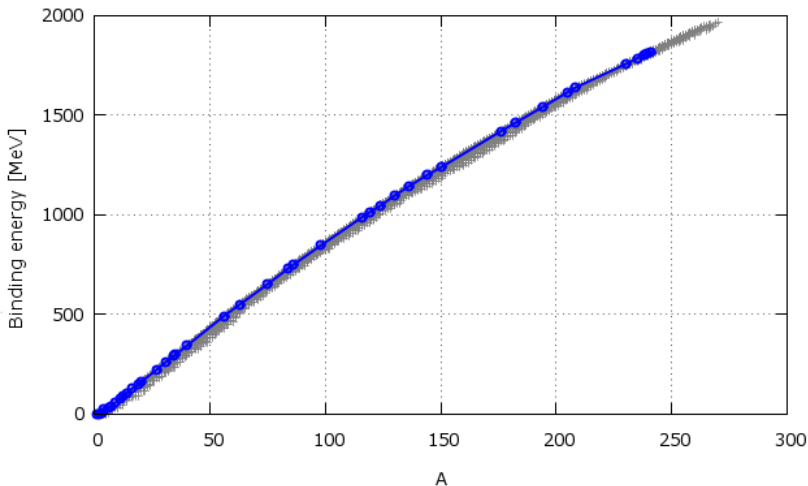




# Binding Energy

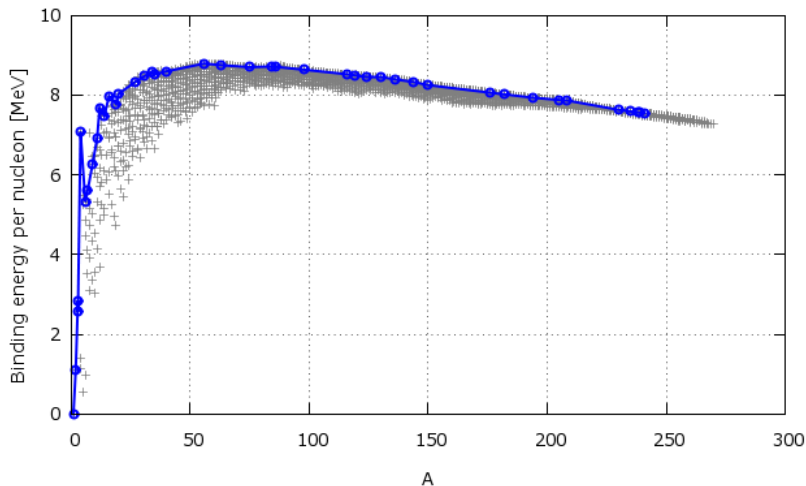
- Mass defect by an atom formation is a result of mass transformation to binding energy
- *Binding energy* is defined as an energy, which must be given to the nucleus to fully disintegrate
- The same binding energy is released by formation of a nucleus from individual components
- The binding energy is usually expressed per one nucleon

# Plot of Binding Energy





# Plot of Binding Energy per Nucleon

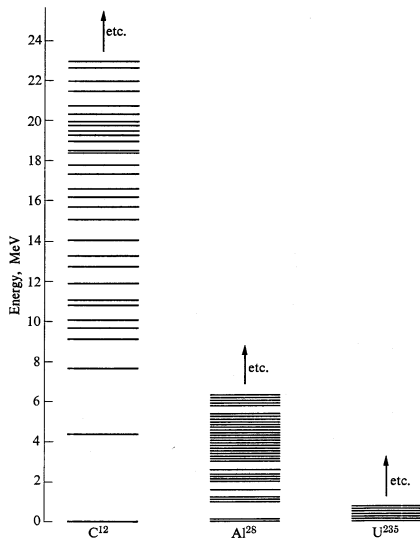




# Excitation States of a Nucleus

- Similar to electrons in atom orbitals, even atom nucleus can be in various energy states
- Energies of excitation states of nuclei are considerably higher
- Energy quantum is released during transition between individual energy levels
- This radiation is called *gamma radiation*
- The excitation energy is distributed among the nucleons in the nucleus. This is opposite to electron orbitals, where first only one particle is excited
- Also inner conversion is possible. In this process, the excitation energy is transferred to the innermost electron, which is in turn released with excitation energy decreased by ionization energy

# Plot of Excitation States of a Nucleus





# Radioactivity

- Atoms found in the nature are either stable or emit particles in order to reach more stable configuration
- Non-stable nuclides have excess or deficiency of neutrons and can undergo beta ( $\beta$ ) decay, or emit alpha ( $\alpha$ ) particles, or neutrons.
- Radioactivity is natural or it can be induced
- During radioactive decay holds four conservation principles:
  - conservation of electric charge (possible neutralisation of positive and negative charge)
  - conservation of mass number (conversion between neutrons and protons is allowed)
  - conservation of mass and energy (sum of mass and energy must be constant)
  - conservation of momentum (distribution of available kinetic energy)



# Radioactive Decay Rates

- Radioactivity is the property of certain nuclides to spontaneously emit particles or gamma radiation
- Radioactive decay occurs at random manner, only average behaviour can be expressed
- There is a certain probability that in a given time a certain fraction of the nuclei within a sample will decay – *radioactive decay constant* ( $\lambda$ )
- The unit of radioactive decay constant is usually  $\text{s}^{-1}$  ( $\text{minute}^{-1}$ ,  $\text{hour}^{-1}$ , etc.)
- *Activity* ( $A$ ) of a sample is the rate of decay of the sample measured in the number of disintegrations that occur per second

$$A = \lambda N, \text{ where } N \text{ is number of atoms in the sample} \quad (1-5)$$



# Units for Measurement of Radioactivity

- There are two common units for measurement of activity – *curie* (Ci) and *becquerel* (Bq)
- $1 \text{ Ci} = 3.7 \times 10^{10} \text{ Bq}$
- Radioactivity decreases exponentially according to the following equation:

$$A = A_0 e^{-\lambda t}, \text{ where } A_0 \text{ is initial activity} \quad (1-6)$$

- According to the same equation decreases also the number of radioactive nuclides

$$N = N_0 e^{-\lambda t}, \text{ where } N_0 \text{ is initial number of nuclides} \quad (1-7)$$





# Radioactive Half-Life

- Radioactive half-life ( $T_{1/2}$ ) is useful for quick estimation of decay of radioactive nuclides
- It is defined as time required for radioactivity to decrease to one-half of its original value
- It is calculated from decay constant as follows:

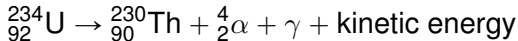
$$T_{1/2} = \frac{\ln 2}{\lambda} \quad (1-8)$$

- After seven half-lives only 0.78 % of the original number of atoms remains
- It is generally assumed that after 10 half-lives, the number of existing atoms is negligible



# Alpha Decay

- Emission of an alpha particle (helium nucleus)
- During decay atomic number is reduced by 2 and mass number is reduced by 4

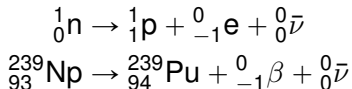


- This decay leads to  ${}_{90}^{230}\text{Th}$  formation and emission of 0.068 MeV gamma radiation
- The kinetic energy corresponds to thorium and  $\alpha$  particle
- The sum of kinetic energy and energy of  $\gamma$  radiation must be equal to mass difference between original uranium nucleus and final particles
- Resulting  $\alpha$  particles carry off 98 % of the kinetic energy

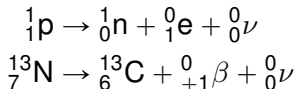


## Beta Decay

- Emission of an electron or a positron of nuclear rather than orbital origin
- Negative electron emission effectively converts a neutron to a proton and an electron



- This example of beta decay is typical for isotopes with excess neutrons
- Typical positron beta decay for nuclei with lack of neutrons follows



- Neutrino has a little practical importance but must be considered for conservation rules



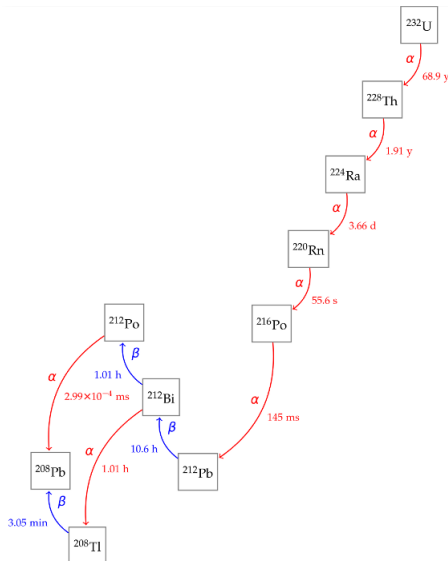
# Gamma Emission

- Gamma radiation is a high-energy electromagnetic radiation that originates in the nucleus
- It is emitted in the form of photons, discrete bundles of energy that have both wave and particle properties
- Daughter nuclides are often left in an excited state after a radioactive decay
- They will drop to the ground state by the emission of gamma radiation



# Decay Chains

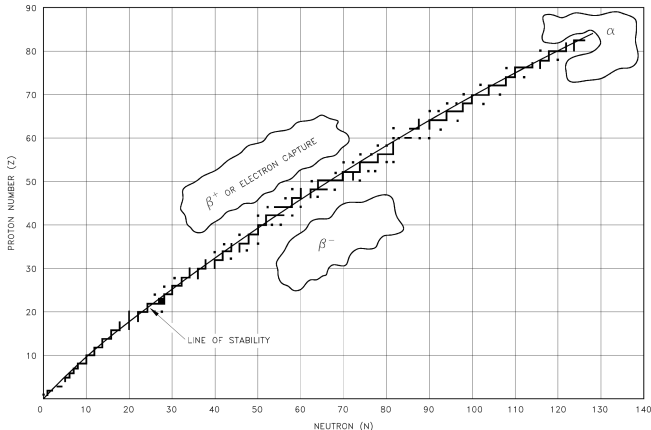
- Daughter nucleus resulting from a decay must not necessarily be stable
- It is often unstable and can undergo further decay
- This is usually true for heavier isotopes
- There is an example for decay chain of  $^{232}\text{U}$  leading to formation of a stable nuclide of  $^{208}\text{Pb}$



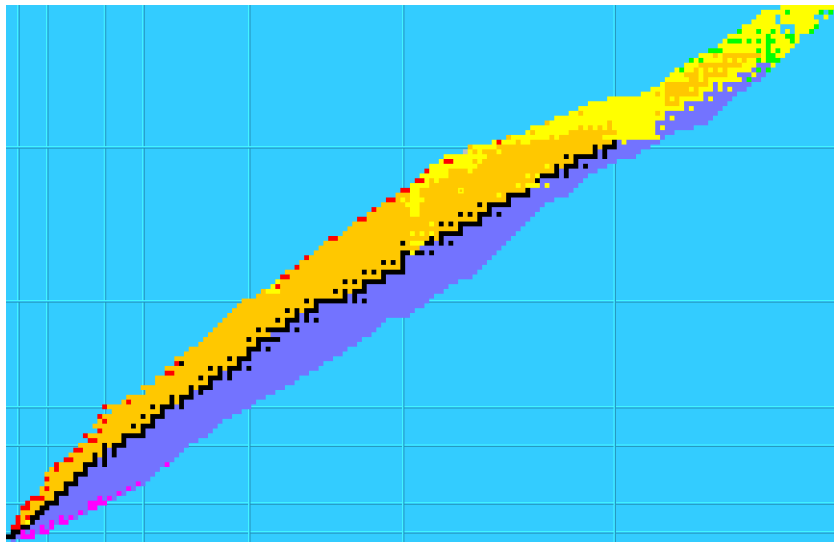


# Predicting Types of Radioactive Decay

- There is a *line of stability* with certain ratio of number of neutrons to number of protons in the nucleus
- Radioactive nuclides tend to decay in a way to get closer to this line



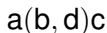
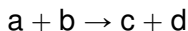
# Line of Stability





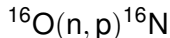
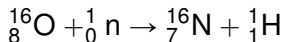
# Nuclear Reactions

- Two nuclear particles can interact together
- As a result two or more nuclear particles or  $\gamma$ -rays are produced
- Nuclear particle can mean nucleus or nucleon



where  $a$  is the target in the rest before reaction and  $b$  is projectile

- For example oxygen bombardment by energetic neutrons:



- All conservation laws valid for radioactive decay have to be preserved also during nuclear reactions