An Introduction to Radioactivity for Engineers

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Foreword

This book is intended as an introduction to the study of radioactivity and to give an idea of its many diverse applications in the field of engineering. An attempt has been made to cover all the relevant areas of the subject, including the types of radiation, the means of detection, the analysis of decay schemes and the necessary safety precautions. To cover all these topics in a book of this size means that the treatment of them must be brief, and for this reason many references are included to books and articles in which the reader will find more detailed information.

The chapters on the industrial and engineering uses of radioisotopes are again intended as a guide to the types of problems where such uses are of great benefit. Again, many references are given to more detailed accounts of such applications.

It is hoped that the text will be suitable for readers engaged in any of the fields of engineering. One difficulty in such an approach is the standardization of units. In general, the units used in the book are those most widely used in the particular application described. For example, reaction cross-sections are quoted in barns, rather than in square metres, and the unit of electron volt is used throughout for radiation energies.

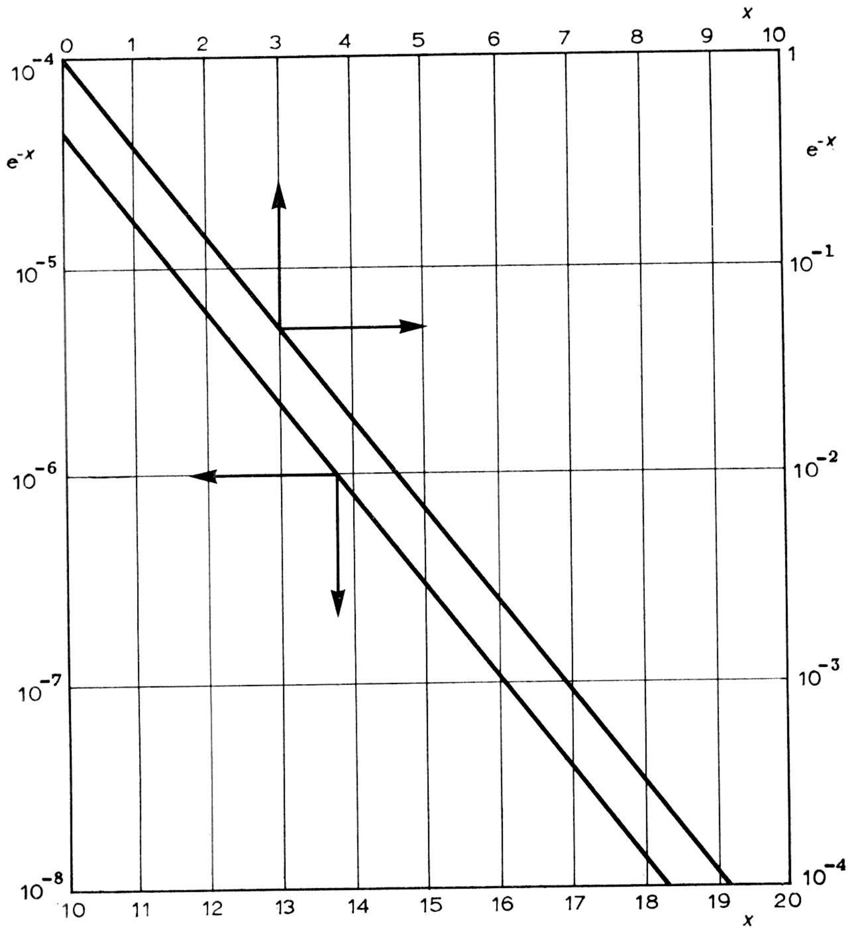
Symbols

- A amount of injected activity
- A area
- A atomic mass number
- A_w atomic weight
- B build-up factor
- c velocity of light
- d deuteron
- **d** distance
- **d** depletion layer thickness in a semiconductor junction
- d_s saturation thickness
- e charge on electron
- e base of Naperian logarithms
- e^- electron
- E energy
- E_m maximum energy in beta particle emission
- f probability of yield from fission reaction
- F calibration constant in total count flow measurement
- h Planck's constant
- I intensity of radiation
- I activity of carbon-14 specimen
- m mass
- M mass of atom
- n number of counts
- n large number of atoms
- n neutron
- N number of neutrons in nucleus
- N impurity concentration in semiconductor
- N total number of counts recorded
- N number of radioactive nuclei in a given sample
- N_A Avogadro's number
- N_b counting rate of background
- N_D number of disintegrations
- N_0 initial number of radioactive nuclei in a given sample
- N_s counting rate of sample
- N_T number of target nuclei per unit volume
- p proton
- p proportion of uranium in a specimen of ore
- P number of protons in a nucleus
- q flow rate of leak
- Q disintegration energy
- Q_1 flow rate in primary circuit
- Q_2 flow rate in secondary circuit
- r distance
- R number of reactions per unit time
- R range in material

R	saturation activity
S	specific activity
t	triton
t	statistical t factor
t	time
t	thickness of target or specimen
t_B	time of count for background
t_s	time of count for sample
T_B	biological half-life
T_{eff}	effective half-life
$T_{\frac{1}{2}}$	radioactive half-life
v	flow velocity
V	volume
V	flow rate
V	voltage across semiconductor junction
W_L	weight of accumulated lead in an ore sample
x	distance
x	number of statistical events
$x_{\frac{1}{2}}$	half-thickness
X	symbol for unknown element
Z	atomic charge
α	alpha particle
α	angle
β	beta particle
β^-	beta particle
β^+	positron
γ	gamma ray
θ	angle
λ	radioactive decay constant
λ	mean free path
λ_i	partial decay constant
μ_L	linear absorption coefficient
μ_M	mass absorption coefficient
ν	frequency
ρ	density
σ	reaction cross-section
σ	standard deviation
σ_f	fission cross-section
Σ	macroscopic reaction cross-section
τ	mean life
ϕ	flux of incident radiation
Ω	solid angle

The symbols, units and Nomenclature used in this book are those recommended by the International Union of Pure and Applied Physics, adopted by its General Council in 1965. These recommendations are in general agreement with those of the following organizations:

1. International Organization for Standardization, Technical Committee I.S.O./T.C.12
2. General Conference on Weights and Measures (1948, 1954, 1960, 1964)
3. International Union of Pure and Applied Chemistry
4. International Electrotechnical Commission, Technical Committees: I.E.C./T.C. 24, 45
5. International Commission on Illumination


x
0 1 2 3 4 5 6 7 8 9 10
10^{-4}
10^{-5}
10^{-6}
10^{-7}
10^{-8}
e^{-x}
1
10^{-1}
10^{-2}
10^{-3}
10^{-4}
e^{-x}
10 11 12 13 14 15 16 17 18 19 20
x

Values of e^{-x}

Miscellaneous Information

Planck's constant, $h = 6{\cdot}6256 \times 10^{-27}$ erg s.
1 barn $= 10^{-24}$ cm^2.
Avogadro's number, $N_A = 6{\cdot}0225 \times 10^{23}$ atoms per gramme atom.
1 u $= 1{\cdot}6575 \times 10^{-24}$ g $\equiv 931{\cdot}459$ MeV.
Mass of electron $= 0{\cdot}0005486$ u.
Energy equivalent of electron mass $= 0{\cdot}51$ MeV.
Charge on electron, $e = 1{\cdot}04 \times 10^{-19}$ coulomb.
1 electron volt $= 1{\cdot}602 \times 10^{-12}$ erg $= 1{\cdot}517 \times 10^{-22}$ Btu.
1 erg $= 10^{-7}$ joule $= 6{\cdot}71 \times 10^{2}$ u.
Energy required to produce an ion-pair in air $= 32{\cdot}5$ eV.
1 curie $= 3{\cdot}7 \times 10^{10}$ d/s $= 2{\cdot}22 \times 10^{9}$ d/min.
1 day $= 8{\cdot}64 \times 10^{4}$ s.
1 week $= 6{\cdot}048 \times 10^{5}$ s.
1 year $= 3{\cdot}1536 \times 10^{7}$ s.
1 röntgen $= 2{\cdot}083 \times 10^{19}$ ion-pairs/cm^3 of air.
1 ångström unit $= 10^{-10}$ m.
1 micron $= 10^{-6}$ m.
1 metre $= 3{\cdot}28$ ft $= 39{\cdot}37$ in.
1 cubic metre $= 35{\cdot}315$ ft^3 $= 1{\cdot}308$ yd^3.
1 cubic centimetre $= 0{\cdot}061$ in^3.
1 litre $= 0{\cdot}22$ gal $= 10^3$ cm^3 $= 0{\cdot}0353$ ft^3.
1 kilogram $= 2{\cdot}679$ lb.

Contents

1 Radioactivity

INTRODUCTION

With almost any engineering method it is possible to use specific techniques without appreciating the underlying physical principles involved. This is especially true of non-destructive and measurement techniques using radioactive isotopes, and doubtless there are many shop floor technicians who are able to make use of particular radioisotopes for certain jobs without too much worry about fundamentals. It is not until the new problem emerges, the difficult measurement has to be made, that the necessity of a fundamental knowledge is fully realized. New techniques and new devices must spring from such a background. With radioisotope techniques in particular, there is an even more potent argument for full understanding. Radioactivity, wrongly used, can be dangerous. It must be handled with care and the full implications of its effects kept in mind at all times.

For these reasons it was decided, when writing this book, to include a brief introduction to the phenomenon of radioactivity and its sources. Thus in the later chapters, when specific techniques and industrial uses are described, the reader will better appreciate the reasons behind the particular choice of parameters.

ATOMIC STRUCTURE

It is assumed that the reader is familiar with the simple concept of atomic structure, as suggested by the Bohr model. In this model the atom is considered as being composed of a central nucleus with a diameter of the order 10^{-12} cm, surrounded by a number of electrons in closed orbits about the nucleus. These orbits have diameters of about 10^{-8} cm. The orbital electrons are grouped in shells by various quantum restraints on the structure. In a consideration of radioactivity, we are not concerned with this extra-nuclear structure, at least for our present purposes. The important point about this model is the electrical neutrality of the atom as a whole.

For our present purposes the nucleus itself can be considered as composed of a number of particles of two distinct kinds. These are the proton, which carries a positive unit charge, $+e$, and the neutron which is uncharged.

Consider that in a particular atom there are Z electrons, each carrying a charge $-e$, orbiting around the nucleus, and that the nucleus is composed of N neutrons and P protons. The condition of electrical neutrality for the atom as a whole yields $Pe-Ze = 0$, i.e. the number of protons in the nucleus is equal to the number of orbital electrons.

The number Z is known as the *atomic charge* or *atomic number* of the atom, and $Z+N$ as the *atomic mass number*, usually denoted by A. The parameters A and Z completely define a particular atomic species, this being known as a *nuclide*.

There have been several re-definitions of mass scales over the years, and quite a bit of confusion over terminology. Nowadays, the scale on which the masses of nuclides are measured is in terms of the *unified atomic mass unit*,* with the symbol u. This is defined as the unit of mass equal to one-twelfth the mass of an atom of carbon of atomic mass number 12. This gives 1 u as $1{\cdot}6575 \times 10^{-24}$ g. On this scale the mass of the neutron is 1·008665 u, the mass on the proton 1·007825 u, and the mass of the electron 0·0005486 u.

From the definition of the mass scale, giving proton and neutron masses of the order unity, it is clear that the atomic mass number will be a whole number approximation to the *nuclidic mass* in u. For example, a nuclide of magnesium which contains 12 protons and 12 neutrons has $A = 24$, and a nuclidic mass of 23·985045 u. The difference between the nuclidic mass and the atomic mass number is called the *mass excess*. A table of mass excesses, based on a value of zero for the carbon-12 nuclide, is given in reference[1].

The chemical properties of the atom, and hence its designation as a particular element, depend upon the number of orbital electrons, i.e. on the atomic number Z. Given Z, the element is uniquely defined. As an example, if a given atom has two orbital electrons it must be helium (assuming that the atom is not ionized or in some similar non-equilibrium

* This replaces the pre-1961 atomic mass unit (amu) which was based on ^{16}O rather than ^{12}C. 1 u = 1·00031792 amu.

state). Similarly an atom with 8 electrons must be oxygen. By increasing Z step by step, it is possible to build up the *periodic table* of elements, given in Table 1.1.

A particular nuclide is denoted by

$$^{A}_{Z}\mathrm{X}$$

where X takes the place of the element symbol. But as Z determines the element, Z and X denote the same thing. Thus the shorthand can be amended to $^{A}\mathrm{X}$. For example, a certain nuclide has six neutrons and six protons. Table 1.1 shows that the element with $Z = 6$ is carbon. Therefore this nuclide can be written $^{12}\mathrm{C}$, or carbon-12.

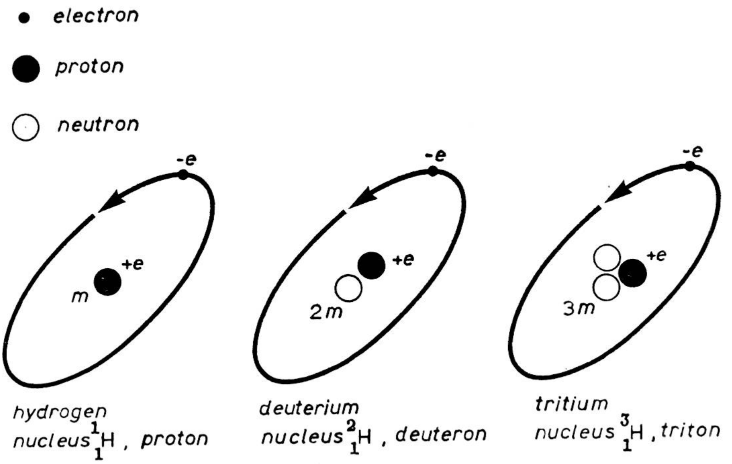


FIG. 1.1 *The three isotopes of hydrogen*

For each element (determined only by Z) there are several nuclides (determined by Z and A) that have the same Z value but different values of A. These different nuclides of the same element are called *isotopes*.

Consider the simplest element, hydrogen with $Z = 1$. Three isotopes are known, with atomic mass numbers of 1, 2 and 3. As Z must remain constant at 1, this means that they have 0, 1 and 2 neutrons respectively. This is illustrated in Fig. 1.1. These isotopes all act chemically as hydrogen, but their nuclidic masses are different. The nuclidic mass of $^{1}\mathrm{H}$ is 1·007825 u, that of $^{2}\mathrm{H}$ (known as *deuterium*) is 2·014102 u, and that of $^{3}\mathrm{H}$ (known as *tritium*) is 3·016049 u. The abundance of deuterium is 0·0156 per cent, and tritium is an artificially produced isotope, not occurring naturally.

The *atomic weight* of an element is defined as the combined nuclidic masses of all the isotopes, weighted according to their natural relative abundances. It is denoted by A_w. In the case of hydrogen it follows that

TABLE 1.1 The periodic table

Atomic charge	*Element*	*Symbol*	*Atomic weight of naturally occurring element*	*Atomic charge*	*Element*	*Symbol*	*Atomic weight of naturally occurring element*
1	hydrogen	H	1·00797	52	tellurium	Te	127·60
2	helium	He	4·0026	53	iodine	I	126·9044
3	lithium	Li	6·939	54	xenon	Xe	131.30
4	beryllium	Be	9·0122	55	caesium	Cs	132·905
5	boron	B	10·811	56	barium	Ba	137·34
6	carbon	C	12·01115	57	lanthanum	La	138·91
7	nitrogen	N	14·0067	58	cerium	Ce	140·12
8	oxygen	O	15·9994	59	praseodymium	Pr	140·907
9	fluorine	F	18·9984	60	neodymium	Nd	144·24
10	neon	Ne	20·183	61	promethium	Pm	(145)
11	sodium	Na	22·9898	62	samarium	Sm	150·35
12	magnesium	Mg	24·312	63	europium	Eu	151·96
13	aluminium	Al	26·9815	64	gadolinium	Gd	157·25
14	silicon	Si	28·086	65	terbium	Tb	158·924
15	phosphorus	P	30·9738	66	dysprosium	Dy	162·50
16	sulphur	S	32·064	67	holmium	Ho	164·930
17	chlorine	Cl	35·453	68	erbium	Er	167·26
18	argon	Ar	39·948	69	thulium	Tm	168·934
19	potassium	K	39·102	70	ytterbium	Yb	173·04
20	calcium	Ca	40·08	71	lutetium	Lu	174·97
21	scandium	Sc	44·956	72	hafnium	Hf	178·49
22	titanium	Ti	47·90	73	tantalum	Ta	180·948
23	vanadium	V	50·942	74	tungsten	W	183·85
24	chromium	Cr	51·996	75	rhenium	Re	186·2

TABLE 1.1 (*continued*)

25	manganese	Mn	54·9380	76	osmium	Os	190·2
26	iron	Fe	55·847	77	iridium	Ir	192·2
27	cobalt	Co	58·9332	78	platinum	Pt	195·09
28	nickel	Ni	58·71	79	gold	Au	196·967
29	copper	Cu	63·54	80	mercury	Hg	200·59
30	zinc	Zn	65·37	81	thallium	Tl	204·37
31	gallium	Ga	69·72	82	lead	Pb	207·19
32	germanium	Ge	72·59	83	bismuth	Bi	208·980
33	arsenic	As	74·9216	84	polonium	Po	210
34	selenium	Se	78·96	85	astatine	At	(211)
35	bromine	Br	79·909	86	radon	Rn	222
36	krypton	Kr	83·80	87	francium	Fr	(223)
37	rubidium	Rb	85·47	88	radium	Ra	(226)
38	strontium	Sr	87·62	89	actinium	Ac	227
39	yttrium	Y	88·905	90	thorium	Th	232·038
40	zirconium	Zr	91·22	91	protoactinium	Pa	231
41	niobium	Nb	92·906	92	uranium	U	238·03
42	molybdenum	Mo	95·94	93	neptunium	Np	(237)
43	technetium	Tc	(99)	94	plutonium	Pu	(242)
44	ruthenium	Ru	101·07	95	americium	Am	(243)
45	rhodium	Rh	102·905	96	curium	Cm	(245)
46	palladium	Pd	106·4	97	berkelium	Bk	(249)
47	silver	Ag	107·870	98	californium	Cf	(249)
48	cadmium	Cd	112·40	99	einsteinium	E	(254)
49	indium	In	114·82	100	fermium	Fm	(252)
50	tin	Sn	118·69	101	mendelevium	Mv	(256)
51	antimony	Sb	121·75				

Numbers in brackets give the atomic mass numbers of the most stable isotope. These values were taken from reference[2]. They are based on C^{12} nuclidic masses.

the atomic weight is

$$1{\cdot}007825\,(0{\cdot}9844) + 2{\cdot}014102\,(0{\cdot}0156) + 3{\cdot}016049\,(0) = 1{\cdot}00797$$

It will be noticed that the masses of the hydrogen isotopes are not obtained by simple addition of neutron masses. For example, the nuclidic mass of ^{1}H plus a neutron is 2·016490 u, yet the mass of deuterium is 2·014102 u. This is a difference of $\Delta m = 0{\cdot}002388$ u, called the *mass defect*. This is because when a proton and a neutron are brought together to form a *deuteron* (the nucleus of deuterium), energy is released in order to bind them together. Conversely energy must be supplied to split them apart. This required energy, the *binding energy*, is obtained from Einstein's equation for the conversion of mass into energy,

$$E = \Delta mc^2 \tag{1.1}$$

where here Δm is the mass defect.

All energies of emitted radiation and particles, as well as the various atomic and nuclear energy levels, are quoted in terms of the *electron volt*, eV. This is the energy that would be acquired by an electron in falling through a potential difference of one volt. From this definition the relationship between other well-known units of energy can be established. In fact,

$$1 \text{ eV} \equiv 1{\cdot}602 \times 10^{-12} \text{ erg} \equiv 1{\cdot}602 \times 10^{-19} \text{ joule} \tag{1.2}$$

For nuclear energy levels, and radiation energies, the electron volt is usually an inconveniently small unit. The units MeV and keV are then used for 10^6 eV and 10^3 eV respectively. Using equation 1.1, with the information that $c = 2{\cdot}99793 \times 10^{10}$ cm/s, and 1 u = 1·0003179 g, then the energy equivalent of 1 u is given by

$$1 \text{ u} \equiv 931{\cdot}459 \text{ MeV} \tag{1.3}$$

In words this means that if say an electron, of mass 0·0005486 u, were completely annihilated, the energy released would be approximately 0·511 MeV. Examples of the use of this relationship are given later.

Example 1.1 *Estimate the atomic weight of naturally occurring magnesium, given that the percentage of each isotope in the natural isotopic mixture is as follows:*

Isotope	^{24}Mg	^{25}Mg	^{26}Mg
Percentage abundance	78·6	10·1	11·3

Atomic weight of magnesium is approximately

$$24\,(0{\cdot}786) + 25\,(0{\cdot}101) + 26\,(0{\cdot}113) = 24{\cdot}32.$$

This is the figure given in Table 1.1. The approximation arises because, as the atomic mass scale is based on carbon, the atomic nuclidic masses are slightly different from the integer values of the atomic mass number. For instance, as seen previously, for magnesium-24, the nuclidic mass is actually 23·985045 u.

PROBLEMS

1.1. What are the elements $^{40}_{18}X$, $^{14}_{7}X$, $^{8}_{4}X$, $^{9}_{4}X$, $^{238}_{92}X$?

1.2 Lithium has two main isotopes, ^{6}Li and ^{7}Li, of relative abundance, 7·4 per cent and 92·6 per cent respectively. What is the atomic weight of the natural isotopic mixture?

1.3. Iron has four naturally occurring isotopes, ^{54}Fe, ^{56}Fe, ^{57}Fe and ^{58}Fe. If the relative abundances of the last three nuclides are 91·52 per cent, 2·245 per cent and 0·33 per cent respectively, and the atomic weight of the natural isotopic mixture is 55·85, what is the percentage abundance of ^{54}Fe?

There are two other terms that are associated with the numbers *A* and *Z*, though not of such general use as nuclide and isotope. Different nuclides having the same value of *A* are called *isobars*, and different nuclides having the same value of $A-Z$ are called *isotones*. This latter definition is the neutron analogue of isotope.

PROBLEMS

1.4. Fill in the blanks, using Table 1.1.

$^{60}_{\dots}Cu$ $\quad$ $^{57}_{28}\dots$ $\quad$ $^{\dots}_{10}Ne$

1.5. State whether the following pairs are isotopes, isobars, or isotones.

^{14}C, ^{14}N; ^{13}N, ^{14}N; ^{15}N, ^{16}O; ^{14}C, ^{15}N; ^{23}Mg, ^{24}Mg

RADIOACTIVITY AND RADIATION

By definition, isotopes have different ratios of neutrons to protons in the nucleus. Some ratios give rise to unstable conditions. This is usually through the neutron to proton ratio being too large. Because of this instability, the nucleus changes its state to attain equilibrium, and in so