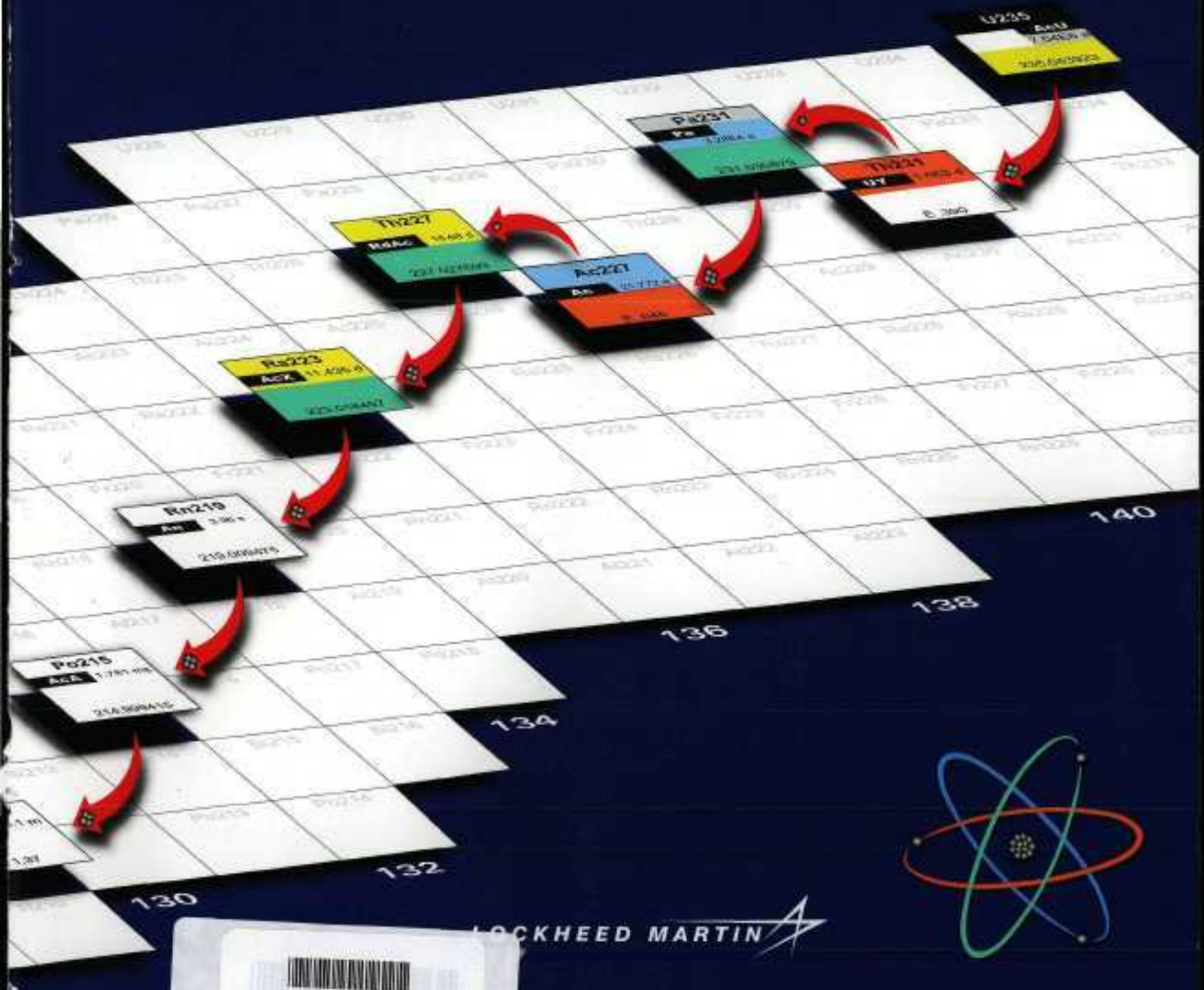
16th edition

Nuclides and Isotopes

Chart of the Nuclides

Wall Chart Information Booklet



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13524

List of Elements by Atomic Number

Z	Sym	Name	Nat. Occ. Isotopes*	Mass (u)	Z	Sym	Name	Nat. Occ. Isotopes*	Mass (u)
1	H	hydrogen	2	1.007 94(7)	59	Pr	praseodymium	1	140.907 65(2)
2	He	helium	2	4.002 602(2)	60	Nd	neodymium	7	144.24(3)
3	Li	lithium	2	6.941(2)	61	Pm	promethium		....
4	Be	beryllium	1	9.012 182(3)	62	Sm	samarium	7	150.36(3)
5	B	boron	2	10.811(7)	63	Eu	europium	2	151.964(1)
6	C	carbon	2	12.0107(8)	64	Gd	gadolinium	7	157.25(3)
7	N	nitrogen	2	14.0067(2)	65	Tb	terbium	1	158.925 34(2)
8	O	oxygen	3	15.9994(3)	66	Dy	dysprosium	7	162.500(1)
9	F	fluorine	1	18.998 4032(5)	67	Ho	holmium	1	164.930 32(2)
10	Ne	neon	3	20.1797(6)	68	Er	erbium	6	167.259(3)
11	Na	sodium	1	22.989 770(2)	69	Tm	thulium	1	168.934 21(2)
12	Mg	magnesium	3	24.3050(6)	70	Yb	ytterbium	7	173.04(3)
13	Al	aluminum	1	26.981 538(2)	71	Lu	lutetium	2	174.967(1)
14	Si	silicon	3	28.0855(3)	72	Hf	hafnium	6	178.49(2)
15	P	phosphorus	1	30.973 761(2)	73	Ta	tantalum	2	180.9479(1)
16	S	sulfur	4	32.065(5)	74	W	tungsten	5	183.84(1)
17	Cl	chlorine	2	35.453(2)	75	Re	rhenium	2	186.207(1)
18	Ar	argon	3	39.948(1)	76	Os	osmium	7	190.23(3)
19	K	potassium	3	39.0983(1)	77	Ir	iridium	2	192.217(3)
20	Ca	calcium	6	40.078(4)	78	Pt	platinum	6	195.078(2)
21	Sc	scandium	1	44.955 910(8)	79	Au	gold	1	196.966 55(2)
22	Ti	titanium	5	47.867(1)	80	Hg	mercury	7	200.59(2)
23	V	vanadium	2	50.9415(1)	81	Tl	thallium	2	204.3833(2)
24	Cr	chromium	4	51.9961(6)	82	Pb	lead	4	207.2(1)
25	Mn	manganese	1	54.938 049(9)	83	Bi	bismuth	1	208.980 38(2)
26	Fe	iron	4	55.845(2)	84	Po	polonium		....
27	Co	cobalt	1	58.933 200(9)	85	At	astatine		....
28	Ni	nickel	5	58.6934(2)	86	Rn	radon		....
29	Cu	copper	2	63.546(3)	87	Fr	francium		....
30	Zn	zinc	5	65.409(4)	88	Ra	radium		....
31	Ga	gallium	2	69.723(1)	89	Ac	actinium		....
32	Ge	germanium	5	72.64(1)	90	Th	thorium	1	232.0381(1)
33	As	arsenic	1	74.921 60(2)	91	Pa	protactinium	1	231.035 88(2)
34	Se	selenium	6	78.96(3)	92	U	uranium	3	238.028 91(3)
35	Br	bromine	2	79.904(1)	93	Np	neptunium		....
36	Kr	krypton	6	83.798(2)	94	Pu	plutonium		....
37	Rb	rubidium	2	85.4678(3)	95	Am	americium		....
38	Sr	strontium	4	87.62(1)	96	Cm	curium		....
39	Y	yttrium	1	88.905 85(2)	97	Bk	berkelium		....
40	Zr	zirconium	5	91.224(2)	98	Cf	californium		....
41	Nb	niobium	1	92.906 38(2)	99	Es	einsteinium		....
42	Mo	molybdenum	7	95.94(2)	100	Fm	fermium		....
43	Tc	technetium		....	101	Md	mendelevium		....
44	Ru	ruthenium	7	101.07(2)	102	No	nobelium		....
45	Rh	rhodium	1	102.905 50(2)	103	Lr	lawrencium		....
46	Pd	palladium	6	106.42(1)	104	Rf	rutherfordium		....
47	Ag	silver	2	107.8682(2)	105	Db	dubnium		....
48	Cd	cadmium	8	112.411(8)	106	Sg	seaborgium		....
49	In	indium	2	114.818(3)	107	Bh	bohrium		....
50	Sn	tin	10	118.710(7)	108	Hs	hassium		....
51	Sb	antimony	2	121.760(1)	109	Mt	meitnerium		....
52	Te	tellurium	8	127.60(3)	110	Ds	darmstadtium		....
53	I	iodine	1	126.904 47(3)	111				
54	Xe	xenon	9	131.293(6)	112				
55	Cs	cesium	1	132.905 45(2)					
56	Ba	barium	7	137.327(7)	114				
57	La	lanthanum	2	138.9055(2)					
58	Ce	cerium	4	140.116(1)	116				

* Number of naturally-occurring isotopes.

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Nuclides and Isotopes

Sixteenth Edition

Revised 2002 by Edward M. Baum, Harold D. Knox and Thomas R. Miller

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This edition of the Chart is dedicated to Frank Feiner, now retired, who contributed to efforts on Charts from 1983 through 1996.

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The Elements and the Periodic Table

Introduction

Man's speculation on the composition of the universe began with the ancient Greeks. In the fifth century B.C., Democritus of Abdera believed that the world was made up of four substances—earth, water, fire and air—and that each of these substances was formed by minute particles that were, themselves, indivisible. These particles were called "atomos" from a Greek word meaning "uncut or indivisible."[1] Qualities—coldness, hotness, wetness, and dryness—were attributed to these classical substances to give them their characteristics. Through the Middle Ages, many variations on these early idealizations were devised.

Robert Boyle gave the modern concept of an element in 1661 in his book *The Sceptical Chymist*.[2] Boyle believed that there existed elemental substances not capable of being broken down into still simpler substances or of being produced by the combination of smaller substances. All other materials were composed of these elemental substances.

Early Developments

Boyle's definition was somewhat vague, however, and chemists of that time could not be sure that some newly developed experimental technique might break down a material previously thought to be an element.[3] Nevertheless, as chemical methods improved, approximately 40 elements were isolated by the early 1800s.

Throughout the Middle Ages, substances were designated by a variety of symbols and letters. Astronomical signs represented the metals known to the ancients: gold (sun), silver (moon), and tin (Jupiter). Two superimposed V's, one inverted, formed the symbol for glass (vitrum in Latin). The symbolic representations of the elements and compounds underwent major changes in the early 1800s. In 1803, John Dalton proposed a more standardized system where circles containing dots, dashes, and the initial letters of the English name of the element designated the atoms of the elements. Initially, this proposal gained some acceptance, but in 1813, the Swedish chemist, Jons Jacob Berzelius, suggested the use of only letter symbols consisting of the first letter of the name of the element and a second letter from the body of the name.[3, 4] This system has remained in effect to the present day.

Between 1803 and 1808, John Dalton made rough determinations of the relative masses of about 35 elements, choosing to make the measurements relative to hydrogen.[5] Berzelius carried out similar work on 41 elements in 1814.[6] These studies seemed to indicate that the relative masses were largely integral. Based on these results, William Prout suggested in 1815 that all elements were built up from the hydrogen atom (Prout's hypothesis[7]). As more precise determinations of the masses were made over the years, Prout's hypothesis was ultimately shown to be incorrect, though for a number of elements, the values were tantalizingly close to integral values.

Following Alessandro Volta's invention[8] in 1800 of the electric pile, or battery, it was soon discovered that elements could be separated by electrolysis, that is, by passing electric currents through solutions containing compounds. In 1834, Michael Faraday showed that a different characteristic weight of each element was produced per unit of charge passed.[9] The charge-carrying species in the electrolysis experiments were called ions. Some elements, such as oxygen, were formed at the positive electrode or anode, and some, such as hydrogen, were formed at the negative electrode, or cathode. Ions of the two different types were called anions and cations, respectively.

As chemical elements were isolated, they were classified according to their chemical and physical properties. Similarities and trends in these properties were noticed. In 1817, Johann W. Dobereiner[10] pointed out the similarities of several groups of three elements (triads), e.g. calcium-strontium-barium and chlorine-bromine-iodine. Several physical properties of a single member of a triad could be predicted by knowing the properties of the other two.

By the mid-1800s approximately 60 elements had been isolated, and new techniques for identifying elements were being developed. One technique was optical spectroscopy. If the light emitted by a material heated to a high temperature was passed through a prism or diffraction grating, that light would be broken up into a spectrum and a number of discrete wavelengths, or lines, could be observed. Robert Wilhelm Bunsen and Gustov Kirchhoff stated that every chemical element was characterized by a unique optical spectrum.[11] Bunsen and Kirchhoff could detect traces of known elements and they predicted new elements could be discovered using this technique. In 1860, in an analysis of a sample of water, they observed two blue lines in the spectrum and attributed these to a new element.[12] The element was named cesium, the name being derived from the Latin word "caesies" meaning "sky blue."[13] During the following year, they detected[14] another element, rubidium. The spectrum of rubidium consisted of two characteristic intense red lines; the name rubidium is from the Latin "rubidus" meaning "deepest red."

Mendeleev's Periodic Table

Demitri Mendeleev began systematizing the properties of the known elements. In 1869, he published a short note entitled "The Correlation between Properties of Elements and Their Atomic Weights."[15] He devised a table, in which the known elements were arranged in columns according to the magnitude of their relative atomic masses, beginning with the smallest. In this vertical form, elements in the same row showed remarkable similarities in their chemical properties, while the properties were periodic from column to column. Originally, Mendeleev called this arrangement the "natural system," but later named it the "periodic system." Yet, there were problems with the arrangement. Certain pairs of elements, e.g. nickel and cobalt, iodine and tellurium, when originally arranged in order of increasing mass, were out of order with their chemical properties. In 1871, Mendeleev revised the table, first, arranging the elements in rows (instead of columns) in order of increasing mass, and repositioning 17 elements from the positions indicated by their masses to new positions indicated by their chemical properties.[16] He argued that the relative mass determinations for these elements must have been in error. The newer table is shown in Figure 1 on page 5.

In order to maintain the periodicity of the properties of the known elements, it was necessary for Mendeleev to leave gaps in the table. He believed these gaps must correspond to as yet

[1] ατομοσ or atomos, from a (not) + τεμνω (cut).
[2] A reproduction of the 2nd Edition of *The Sceptical Chymist*, originally published in 1680, has been published by The Classics of Science Library, a Division of Gryphon Editions, 1997, New York.
[3] Isaac Asimov, *Understanding Physics, Volume III, The Electron, Proton and Neutron*, (The New American Library, Inc., New York, April 1969).
[4] A more complete description of the work of Dalton and Berzelius on chemical symbols and on mass determinations can be found in F.P. Armitage, *A History of Chemistry*, (Longmans, Green and Company, London, 1920).
[5] J. Dalton, *A New System of Chemical Philosophy*, (Manchester, 1808), p. 213.
[6] J.J. Berzelius, Ann. Phil. **3** 353 (1814).
[7] W. Prout, Ann. Phil. **6** 269 (1815).
[8] A. Volta, Phil. Trans. R. Soc. **90** 430 (1800).
[9] M. Faraday, Phil. Trans. R. Soc. **124** 77 (1834).
[10] J.W. Dobereiner, Ann. Physik (Gilbert) **57** 435 (1817).
[11] R. Bunsen and G. Kirchoff, Phil. Mag. **20** 89 (1860).
[12] R.W. Bunsen and G. Kirchhoff, Chem. News **2** 281 (1860).
[13] N.E. Holden, "History of the Origin of the Chemical Elements and Their Discoverers", BNL-NCS-63850-01/10-REV, Brookhaven National Laboratory, 2001. On the Internet, see http://www.nndc.bnl.gov/nndc/history/.
[14] R.W. Bunsen, Chem. News **3** 357 (1861).
[15] D. Mendeleev, Zeits. Fur Chemie **12** 405 (1869).
[16] D. Mendeleev, Annal. Der Chemie Supp. VIII 133 (1871).

Reihen	Gruppe I – R^2O	Gruppe II – RO	Gruppe III – R^2O^3	Gruppe IV RH^4 RO^2	Gruppe V RH^3 R^2O^5	Gruppe VI RH^2 RO^3	Gruppe VII RH R^2O^7	Gruppe VIII – RO^4
1	H=1							
2	Li=7	Be=9.4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27.3	Si=28	P=31	S=32	Cl=35.5	
4	K=39	Ca=40	– =44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59 Ni=59, Cu=63.
5	(Cu=63)	Zn=65	– =68	– =72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	– =100	Ru=104, Rh=104 Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

Figure 1. Mendeleev's Periodic Table of 1871 (Ref. 16). *The known masses are given after the equal signs. Question marks indicate uncertainty as to the "correct" location of the elements in the table. Note that copper (Cu), silver (Ag), and gold (Au) were provisionally included in both Groups I and VIII.*

undiscovered elements. He predicted the existence and properties of gallium, scandium and germanium. All three of these elements were discovered within the next 15 years and their properties were found to be as he had predicted.

By the early 1890's approximately 80 elements had been isolated, including uranium, thorium, and most of the elements known today through bismuth. Though some gaps still remained, Mendeleev's table nicely showed the periodic nature of the properties of the known elements. When these gaps were filled, would the table be complete, or was it possible that whole families of as yet undiscovered elements might exist? The latter proved to be the case. Between 1894 and 1898, the inert gases—helium, neon, argon, krypton, and xenon—were discovered.[13] As the relative masses and chemical properties of these gases were determined, a new previously unexpected column was inserted into the periodic table. Until the cause of the periodicity of the properties of the elements was understood, there could be no guarantee that still other unknown families of elements might not exist.

The Rare Earths

Certain metallic oxides, such as alumina, silica and zirconia, which were found in minerals, were called "earths." These oxides were difficult to reduce and, because of this difficulty, some had been erroneously classified as elements.[17] In 1794, Johan Gadolin isolated a new "earth," later named yttria, from minerals found near Ytterby, a village in Sweden.[18] A second new earth, ceria, was isolated independently by Martin H. Klaproth[19] and by Jons J. Berzelius and W. Hisinger[20] in 1803. Initially, it was believed that only a few geographically isolated mineral deposits contained these oxides, and they were, therefore, called the "rare earths." Over the course of the next hundred years, the two materials were found to contain sixteen new elements. In yttria, the elements scandium, yttrium, gadolinium, terbium, dysprosium, holmium, erbium, thulium, ytterbium, and lutetium were identified. Lanthanum, cerium, praesodymium, neodymium, samarium and europium were found in ceria.

The location of these elements in the periodic table was uncertain. In 1871, Mendeleev had included four known rare earths, Di (didymium, later shown not to be an element), Ce (cerium), Er (erbium), and La (lanthanum) in his table. These four materials were placed in two different rows or series.

As more of these elements were isolated, their chemical properties were found to be similar and their atomic masses were closely spaced. In 1905, A. Werner published a new version of the periodic table including the known rare earths and the known inert gases.[21] This "long form" of the periodic table is shown in Figure 2 on page 6. The twelve series in Mendeleev's table were combined into seven periods in Werner's table. Gaps left in periods (rows) 4 and 5 of Werner's table provided space for the rare earths in period 6.

The Electron

The conduction of electricity through rarified gases was a popular subject of study by scientists in the second half of the nineteenth century. Sir William Crookes used a partially evacuated glass tube, known as a discharge tube, containing two electrodes attached to an external voltage source to carry out such investigations.[22] Crookes observed that if the gas pressure in the discharge tube was reduced to some critical value, long bright streamers extended from the negative electrode (cathode) along the length of the tube. These streamers were called cathode rays. Transverse electric and magnetic fields could deflect the rays, and the direction in which they were deflected indicated that the charge of the particles comprising the rays was negative.

In 1897 at the Cavendish Laboratory in Cambridge, Joseph J. Thompson made measurements of the charge-to-mass ratio (e/m) of the cathode ray particles using electric and magnetic fields to deflect the beam.[23] To be certain that the properties of the cathode rays were independent of the materials used, several measurements were carried out using different gases in the discharge tube and different electrode materials. All measurements gave the same, surprisingly large, value for the ratio.

Prior to Thompson's e/m measurement of the cathode rays, values of the charge-to-mass ratio for some ionized atoms had been obtained in experiments involving electrolysis of solutions. The largest ratio was obtained for an ionized hydrogen atom.[24] Thompson's e/m value for the cathode ray particles was 1836 times larger than that of hydrogen. Either the charge of the cathode ray particle was much larger than that of the ionized hydrogen atom, or the particle mass was much smaller.

[17] See James F. Spencer, *The Metals of the Rare Earths*, (Longmans, Green and Co. London, 1919), for a detailed account of the discovery of the rare earths as well as changes in the periodic table during this time.

[18] J. Gadolin, K. Svenska. Vet-Ak. Hdl. Ordv. **15** 137 (1794).

[19] M.H. Klaproth, Crell's Ann. **1** 307 (1801).

[20] J.J. Berzelius and W. Hisinger, Gehl. Ann. **2** 303 (1804); **2** 397 (1804).

[21] A. Werner, Ber. Dtsch. Chem. Ges. **38** 914 (1905).

[22] W. Crookes, Phil. Trans. Part I, 135 (1879), see W.F. Magie, *A Source Book in Physics*, (McGraw Hill, New York, 1935).

[23] J.J. Thompson, Phil. Mag. **44** 293 (1897).

[24] In 1920, E. Rutherford suggested the name "proton" (from the Greek "πρωτον", meaning "first") for the ionized hydrogen atom. See F.W. Aston, *Mass Spectra and Isotopes, Second Edition*, (Longmans, Green & Company, NY, 1941), p. 192.

1	H 1.008																																He 4
2	Li 7.03																									Be 9.1	B 11	C 12	N 14.04	O 16.00	Fl 19	Ne 20
3	Na 23.00																									Mg 24.36	Al 27.1	Si 28.4	P 31.0	S 32.06	Cl 35.46	A 39.9
4	K 39.10	Ca 40.1																Sc 44.1	Ti 48.1	V 51.2	Cr 52.1	Mn 55.0	Fe 55.9	Co 58.9	Ni 58.7	Cu 63.6	Zn 65.4	Ga 70	Ge 72	As 75.0	Se 79.1	Br 79.96	Kr 81.12
5	Rb 85.4	Sr 87.6																Y 89.0	Zr 90.7	Nb 94	Mo 96.0	–	Ru 101.7	Rh 103.0	Pd 106	Ag 107.9	Cd 112.4	In 114	Sn 118.5	Sb 120	Te 127.6	J 126.9	X 128
6	Cs 133	Ba 137.4	La 138	Ce 140	Nd 143.6	Pr 140.5	–	–	Sa 150.3	Eu 151.8	Gd 156	Tb 160	Ho 162	Er 166	–	Tu 171	Yb 173.0	–	–	Ta 183	W 184.4	–	Os 191	Ir 193.0	Pt 194.8	Au 197.2	Hg 200.3	Tl 204.1	Pb 206.9	Bi 208.9	–	–	–
7	–	Ra 226	Lac 7	Th 232.8	–	–	–	–	–	U 238.5	–	–	–	–	–	Ac 7	–	–	–	–	–	–	–	–	–	–	–	–	Pbo ?	Bio ?	Teo ?	–	–

Figure 2. The Long Form of the Periodic Table Published by A. Werner in 1905 (see Ref. 21). *The dark lines around certain pairs of elements indicate the ordering of the pair was based on chemical properties, not atomic masses.*

Thompson believed the latter and that the negative charge on the particle was equal in magnitude to that of a singly ionized atom. Robert A. Millikan would later confirm this in experiments carried out in 1910.[25] The cathode ray particles were called electrons, a name originally suggested by the British scientist G. Stoney in 1894 for the charge of a univalent ion.[26]

Optical Spectroscopy

Anders Angstrom initially investigated the optical spectrum of hydrogen in 1855.[27] In the visible region, nine lines were observed. In 1885, Johann Balmer, a Swiss mathematician, showed that the wavelengths, λ, of these lines could be expressed by the formula

$$\lambda = B\frac{n^2}{n^2-4}$$

where B was a constant and n an integer that could take on values from 3 to 11.[28] Upon learning that additional, shorter wavelength lines of this series had been observed in astronomical studies, Balmer showed that the above equation also correctly predicted the wavelengths of these newly observed lines.

Balmer's work prompted additional searches for empirical formulae that would give the wavelengths of spectral lines of other elements. Johannes Rydberg did much of this work.[29] Rydberg found it more convenient to deal with the wave number, k, and rewrote Balmer's formula as

$$k = \frac{1}{\lambda} = R_H\left(\frac{1}{2^2} - \frac{1}{n^2}\right), \quad n = 3, 4, 5, \ldots$$

where R_H is the Rydberg constant for hydrogen. The value of R_H was empirically determined. (The Bohr model of the hydrogen atom, page 9, would later provide an expression for the Rydberg constant in terms of fundamental physical constants.) Formulae of this type, but generally more complex, were obtained for a number of other series observed in the spectra of heavier elements.

A Model of the Atom

The proof of the independent existence of the electron was a starting point for theories of atomic structure. In 1910 Thompson proposed one of the first models of the atom.[30] His "plum pudding" model of internal structure depicted the atom as a homogeneous sphere of positive fluid, the "pudding," in which were imbedded the negatively charged electrons, the "plums." In this model, the electrons, which repel each other but are attracted to the positive charge, assume certain stable positions inside the atom. If an external force, e.g. the violent collisions between atoms in a hot gas, disturbed the electron distribution, the electrons would vibrate about their equilibrium positions and emit electromagnetic radiation. Thompson's model, then, provided a qualitative picture for the optical and X-ray emission spectra from excited atoms. Calculations of the spectrum of radiation emitted from this model, however, were not in quantitative agreement with experimental data. Conclusive proof of the inadequacy of this model would come in 1911 from experiments, described on page 8, carried out by two students working under Ernest Rutherford.

X-rays

In 1895, while experimenting with electric current flow in a discharge tube, Wilhelm Konrad Röentgen noticed that salts of barium would fluoresce when brought near the tube.[31] He determined that whatever caused the fluorescence in the salts originated at the point of the discharge tube where the cathode rays struck the glass wall. Apparently, some type of radiation was produced when the cathode rays impacted the end of the tube. Röentgen called this radiation X-rays, because of its unknown nature. He found the radiation to be energetic—it could pass through many layers of paper and even thin layers of metal. If metallic anodes were used in the cathode ray tube, even more X-rays were produced; anodes made of heavy metals produced more X-rays than anodes of light metals. The X-rays traveled in straight lines and were not affected by electric or magnetic fields. They were found to ionize gases, making the gases conductors of electricity. In 1912, experiments involving the scattering of X-rays from crystals indicated the rays were wave-like in character, i.e. a form of electromagnetic radiation.[32]

Radioactivity

After the discovery of X-rays, scientists wondered if this radiation might also be emitted by certain materials that were known to be fluorescent—materials that would absorb light of a particular wavelength and then give off light of a longer wavelength. In 1896, Antoine-Henri Becquerel began such

[25] R.A. Millikan, Phil. Mag. **19** 209 (1910).
[26] G.J. Stoney, Phil. Mag. **38** 418 (1894).
[27] A.J. Angstrom, Phil. Mag. **9** 327 (1855).
[28] J.J. Balmer, Ann. der Physik und Chemie **25** 80 (1885). This reference is a summary of two earlier publications by Balmer on the hydrogen spectrum.
[29] J.R. Rydberg, K. Svenska Vet-Ak. Hall. Oefv. **23** 1889.
[30] J.J. Thompson, Proc. Cambridge Phil. Soc. **15**, 465 (1910).
[31] W.K. Röentgen, Sitzber. Wurzberger Phys.-Med. Ges., Dec. 1895; Nature **53** 274 (1896).
[32] W. Friedrich, P. Knipping and M. von Laue, Ber. Bayer. Akad. Wiss. 303 (1912).

studies on a uranium-containing compound, potassium uranyl sulfate.[33] He found that after exposing this material to sunlight, radiation from the compound would penetrate paper and darken a photographic plate. More surprisingly, he found the material would darken the plate without having been exposed to sunlight. This would indicate that the material must have been emitting some type of radiation all the time. Becquerel found this radiation to be as penetrating as X-rays and capable of ionizing gases. Two years later, Marie Curie observed a similar radiation emitted from compounds containing thorium.[34]

In 1898, Marie Curie began detailed studies consisting of standard chemical separations of the materials found in the uranium ore, pitchblende. The intensities of the radiation emitted by the separated compounds changed as a function of time and measurements of the activities, or rates of decay, were carried out. The activities were measured with an ionization chamber, first developed[35] by the Curies. In the course of these studies, Mme. Curie isolated a new element, chemically similar to bismuth. She named the element polonium, in honor of her native country, Poland. In a July 1898 publication announcing this discovery, the word "radio-active" was used for the first time in describing the materials in the uranium compounds.[36] The Curies continued their studies and, in December 1898, announced the discovery of another element, which they named radium, from the Latin word "radius," meaning, "ray."[37] The following year, Andre Debierne, a former student of Pierre Curie, discovered a third element, actinium (from the Greek "atkis," meaning "beam or ray"), in similar studies.[38]

Ernest Rutherford began studying the uranium and thorium compounds in 1898. He determined that two types of radiation were being emitted from the materials, "one that is very readily absorbed, which will be termed for convenience α radiation, and the other of a more penetrative character, which will be termed the β radiation."[39] In 1899, F. Giesel determined that β-rays were the same as cathode rays.[40] In 1900, Paul Villard identified still a third type of radiation, which was designated gamma radiation.[41] This was subsequently shown to be a form of electromagnetic radiation, like light and X-rays, only having higher energy. In 1904, Rutherford, by collecting the α-rays in an evacuated tube and detecting the buildup of helium, demonstrated that α-particles were doubly ionized helium atoms.[42]

Ernest Rutherford and Frederick Soddy quantitatively studied the activities of the materials chemically separated from the uranium and thorium compounds.[43] They found an exponential decay rate for the activity, A, as a function of time, t,

$$A(t) = A_0 e^{-\lambda_d t}$$

where λ_d is the disintegration constant[44] and A_0 is the initial activity. Another related constant that was used to characterize the materials was the half-life, $t_{1/2}$, the time needed for half of the radioactive atoms in a given sample to disintegrate.

$$t_{1/2} = \ln(2)/\lambda_d = 0.693/\lambda_d$$

The various "daughter" products in the series could then be characterized by their half-lives as well as the types of radiation emitted. In 1903, Rutherford and Soddy published a paper, "Radioactive Change," describing the change of an atom of one material into an atom of another material by the emission of an alpha or beta ray.[45] An atom of uranium would spontaneously emit an alpha ray, changing into a chemically distinct atom of another material called uranium X_1. This atom would then emit a beta ray, changing into an atom of another material, called uranium X_2, and so on. A series of radioactive materials was thus generated.

Actually, three naturally-occurring radioactive decay series—the thorium, uranium, and actinium series—were being studied during this time (see Figure 12, page 25). Names were given to more than 40 radioactive "species" isolated in these series.[46] In the uranium series, there were uranium I (UI), uranium X_1 (UX_1), uranium X_2 (UX_2), ionium (Io), radium A (RaA), radium B (RaB), and others. In the thorium series, there were mesothorium I ($MsTh_1$), mesothorium II ($MsTh_2$), radiothorium (RdTh), thorium X (ThX), thorium A (ThA), thorium B (ThB), and others. Between 1906 and 1907, several investigators showed that ionium, a uranium decay product, and radiothorium, a thorium decay product, when mixed with thorium, could not be separated by any chemical processes. The three substances had different radioactive properties, but were chemically the same. In addition, A.S. Russell and R. Rossi showed that the optical spectra of pure thorium and of a mixture of thorium and ionium were indistinguishable.[47] It would seem that these substances should occupy the same position in the periodic table. Soddy introduced the term "isotope" to describe these materials.[48] That a given element could exist in more than one chemically identical form would explain the large number of materials that were being observed in the studies of the radioactive decay chains.

Confirmation of the existence of isotopes came in 1911 when Soddy published a comparison of the atomic masses of the stable element lead obtained from uranium rich ores and from thorium rich ores.[49] The lead from the uranium ores had an average atomic mass slightly less than that from the thorium ores. More conclusive proof of isotopes came, when, first, J.J. Thompson[50] and, then, Francis William Aston[51] carried out experiments consisting of passing positively charged neon ions through a magnetic field onto a photographic plate. Aston found two spots, one ten times as large as the other, on the plate. He concluded that the neon atoms could have either of two masses, and that the abundance of the lighter mass was ten times that of the heavier. (A third isotope of neon would be discovered[52] later.) Modifying Prout's hypothesis, Aston proposed the whole number rule stating that the fractional atomic masses of elements are due to the presence of two or more isotopes, each of which has an approximately integral mass.[53]

Van den Broek had suggested that the number of elementary positive charges on any nucleus corresponds to the ordinal number of the element in the periodic table.[54] In 1913, Soddy announced the displacement laws of radioactive decay.[55] He proposed that the result of alpha emission (alpha particles have a +2 electric charge) from a material would be the formation of a "daughter" product that is two places to the left of the "parent" in the periodic table. Furthermore, the emission of a beta particle, having a charge of minus one, would result in the formation of a

[33] A.H. Becquerel, Compt. Rend. **122** 420 (1896); **122** 501 (1896).
[34] M. Sklodowska, C.R. Acad. Sci., Paris **126** 1101 (1898).
[35] see S. Quinn, *Marie Curie: a life*, (Simon & Schuster, New York, NY, 1995), p. 145.
[36] P. Curie and M. Sklodowska Curie, Compt. Rend. **127** 175 (1898).
[37] P. Curie, Mme. P. Curie, and G. Bemont, Compt. Rend. **127** 1215 (1898).
[38] A. Debierne, Compt. Rend. **130** 960 (1900).
[39] E. Rutherford, Phil. Mag. **47** 109 (1899).
[40] F. Giesel, Ann. d. Phys. **69** 834 (1899).
[41] P. Villard, Compt. Rend. **130** 1178 (1900).
[42] E. Rutherford and T.D. Royds, Phil. Mag. **17** 281 (1909).
[43] E. Rutherford and F. Soddy, Phil. Mag. **4** 370 (1902); **4** 569 (1902); **5** 576 (1903).
[44] Due to the use of the symbol "λ" for wavelength previously, the subscript "d" is added when denoting the disintegration constant; this is not the standard notation.
[45] E. Rutherfied and F. Soddy, Phil. Mag. **5** 576 (1903).
[46] A system for naming the radioactive "elements" was suggested by E. Rutherfied but was not uniformly used. See N.R. Campbell, Nature **76** 638 (1907).
[47] A.S. Russell and R. Rossi, Proc. Roy. Soc. **87** 478 (1912).
[48] The word "isotope" is from the Greek words ἴσος, meaning equal, and τόπος, meaning place and was suggested by Soddy since isotopes occupy the same place in the periodic table of the elements. See F. Soddy, Nature **92** 399 (1913).
[49] F. Soddy, Trans. Chem. Soc. **99** 72 (1911).
[50] see F.W. Aston, *Isotopes*, (Edward Arnold & Co., London, 1923), pp. 33-34.
[51] F.W. Aston, Phil. Mag. **39** 449 (1920).
[52] T.R. Hogness and H.M. Kvalnes, Phys. Rev. **32** 942 (1928).
[53] F.W. Aston, Proc. Roy. Soc. **A115** 487 (1927).
[54] A. Van den Broek, Physik. Z., **14** 32 (1913).
[55] F. Soddy, *Chemistry of the Radio-Elements*, (Longmans, Green & Co., Ltd, London, 1914).

"daughter" product that is one place to the right of the "parent" on the periodic table.

Rutherford's Model of the Atom

Beginning in 1906, Rutherford carried out a series of experiments with a beam of high-energy alpha particles bombarding a very thin metal foil, detecting the number of alphas scattered at various angles from the target. With the Thompson plum pudding model of the atom, in which the positive charge was uniformly distributed throughout the atom, a sufficiently large concentration of charge would not exist in any one region of the atom to affect the alpha particles significantly. One would expect the alpha particles to pass through the material nearly undeflected. Rutherford did find most of the alpha particles passed straight through the foil or were scattered only slightly from their original paths. However, a small percentage was significantly deflected with some reversing their directions. This was totally unexpected. Rutherford thought "it (the experimental result) was about as credible as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you."[56]

To explain these results, in 1911, Rutherford postulated that the atom does not consist of a uniform sphere of positive electrification, but that the positive charge is concentrated in a very small region, called the nucleus, at the center of the atom.[57] In his dynamic planetary model, the nucleus plays the role of the sun and the electrons correspond to the individual planets of the solar system, revolving about the sun. This model, along with the classical laws of electricity and mechanics, provided an adequate explanation of alpha particle scattering. Subsequent experiments performed on seven different scattering materials and for different alpha energies verified Rutherford's theory.

While the model satisfactorily explained the results of alpha particle scattering, there were still questions about the stability of such an atom. Electromagnetic theory demands that an oscillating or revolving electric charge emit electromagnetic radiation. Such emission would result in the loss of energy by the emitting particle. Applied to orbital electrons, this energy loss would cause a steady contraction of the system, since the electrons would spiral into the central nucleus as their rotational energy was dissipated. This process would occur very rapidly and would directly contradict the permanent existence of atoms. Also, if the radiation pattern produced by the atom were related to the energy radiated by the moving electron, this radiant energy would change with the radius of curvature of the electron's path. The pattern would consist of a continuous range of wavelengths. However, this is contradicted by experiments where individual, well-defined lines characterize the optical (and X-ray) spectra for a given element, not a continuous range of wavelengths.

Quantum Theory

In the early 1900s, it became obvious that not all physical phenomena could be adequately explained by the laws of classical physics. The thermal radiation spectrum emitted by a "blackbody" (a perfectly absorbing body) was one such phenomenon. A cavity inside a blackbody heated to a high temperature would contain electromagnetic radiation, generated by vibrating particles in the walls of the cavity. The distribution of the wavelengths of the radiation depended only on the temperature, not on the composition of the material.

Early attempts to explain the observed wavelength distribution failed. Theoretical predictions by Rayleigh[58] and Jeans[59] using classical statistical mechanics agreed with the experimentally measured distribution for long wavelengths, but approached infinity as the wavelength became small. A comparison of the predictions of Rayleigh-Jeans with

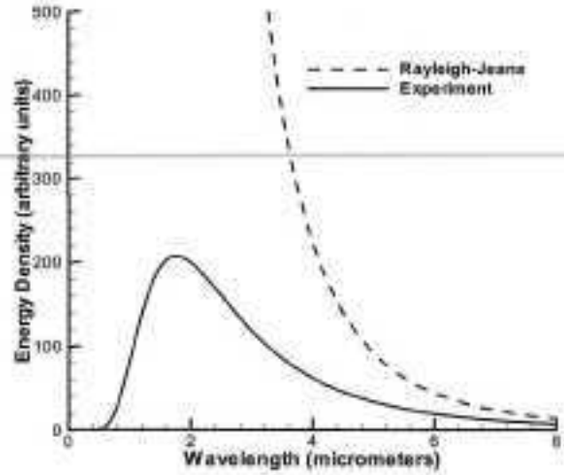

Figure 3. Comparison of the Rayleigh-Jeans Prediction for the Spectrum of Blackbody Radiation with Experimental Measurements (T = 1640 K). *At long wavelengths, the theoretical predictions agree with experimental results.*

experimental measurements is shown in Figure 3. In 1900, Max Planck[60] resolved this problem by making the radical assumption that the particles in the cavity wall could radiate energy only when the total energy was a multiple of some discrete amount (or quantum)[61] of energy. The energies were given by

$$E = nh\nu$$

where n is an integer, ν is the frequency of the emitted electromagnetic radiation, and h is a constant (later named Planck's constant). With this assumption, the predicted spectrum agreed very well with experimental data.

Classical physics also failed in explaining the photoelectric effect. When light (electromagnetic radiation) is incident on a metallic surface, electrons are sometimes emitted from the metal. According to classical physics, the energy associated with an electromagnetic wave is proportional to the square of the amplitude or intensity. Accordingly, the emission of the photoelectrons should occur when the intensity of the light is sufficient to supply the necessary energy to eject the electrons. This is not the case.

Experimentally, it is found that the production of photoelectrons depends on the frequency of the incident light. Below a certain "threshold" frequency, which depends on the metal, no electrons are emitted regardless of the intensity of the light. Above this frequency, electrons are emitted and their energy is proportional to the frequency of the light. For light of a given frequency, the number of photoelectrons depends on the intensity of the light.

In 1905, Albert Einstein proposed a mechanism for the emission of photoelectrons from a metal.[62] He assumed that light has particle-like characteristics and that the energy, E, associated with these quanta was given by

$$E = h\nu$$

where ν is the frequency of the light and h is Planck's constant. A quantum of light would interact with an electron in the metal. If the energy imparted to the electron were sufficiently large, the electron would be emitted from the metal. The maximum kinetic

[56] Quoted in J. Needham and W. Pagel, Ed., *Background to Modern Science*, (Cambridge University Press, 1940), p. 68.
[57] E. Rutherford, Phil. Mag. **21** 669 (1911).
[58] Lord Rayleigh, Phil. Mag. **49** 539 (1900).
[59] J.H. Jeans, Phil. Mag. **10** 91 (1905).

[60] M. Planck, Verh. D. deut. Physik. Gesell **2** 237 (1900); Ann. Physik **4** 553 (1901).
[61] From the Latin word "quantus", meaning "how great".
[62] A. Einstein, Ann. Physik **17** 132 (1905).

energy of the electron is given by

$$T_{max} = h(\nu - \nu_0)$$

where ν_0 is a constant, the threshold frequency, characteristic of the metal. Experimental measurements confirmed the validity of this equation; higher precision measurements, made later, were used to accurately determine the value of Planck's constant.

Relativity

In the early 1900s, Einstein also resolved some inconsistencies between the behavior of electromagnetic radiation and the laws of classical physics.[63] According to his special theory of relativity, published in 1905, the mass of an object in a given frame of reference is given by

$$m = m_0(1 - \frac{v^2}{c^2})^{-1/2}$$

where m_0 is its mass when at rest, v is its velocity in the reference frame, and c is the speed of light in a vacuum. No particle with a finite rest mass can travel at v = c in any reference frame, because its mass and energy would be infinite. Furthermore, mass and energy are equivalent, the total energy content of an object of mass m being given by the equation $E = mc^2$. (Beta particles emitted in radioactive decay can have velocities close enough to c that relativistic effects become important. Heavier particles of comparable energy travel much more slowly, so that their behavior is adequately described by classical physics.)

Studies of X-Ray Spectra

Charles W. Barkla continued studies of Röentgen's X-rays. In 1911, he showed that the energy of the X-rays emitted from a cathode ray tube depended on the anode material.[64] The X-ray spectra consisted of a continuous broad band of wavelengths on which groups of a few sharp, intense lines were superimposed. The "hardest" or most energetic group of lines was called the K-series, the next, the L-series, then the M-series, and so on.[65]

In 1913 Henry G.J. Moseley studied the K-series spectra for the elements from calcium to zinc and the following year looked at the L-series spectra for elements up to gold.[66] Generally, as the atomic mass increased, the frequency of the X-ray lines increased, but there were irregular variations. There was, however, a regularity of the frequency as a function of the position of the element in the periodic table. If each element was assigned a number, Z, according to its position on the periodic table, Moseley found the square root of the X-ray frequency was proportional to Z-1 for each element (see Figure 8, page 16). He reasoned that Z (now called the atomic number) must correspond to the positive charge on the nucleus. He concluded the ordering of elements in the periodic table should be based on Z, not on atomic masses.

This insight into the ordering of the elements in the periodic table was an important advance and provided a new tool to scientists. The proper order of the known elements, including the rare earths, was now understood. The elements still missing from the periodic table could now be predicted with certainty. In 1913 the heaviest element known was uranium (Z = 92). Below that, seven gaps still existed, corresponding to Z = 43, 61, 72, 75, 85, 87 and 91. These were the only elements yet to be isolated. There were no more missing families of elements, as the inert gases had been earlier. Figure 4 (page 10) shows the status of the periodic table at that time.

[63] A. Einstein, Ann. Phys. **17** 891 (1905).
[64] C.W. Barkla, Phil. Mag. **22** 396 (1911).
[65] *Ibid.* These series had earlier been designated by the letters A, B, etc. Barkla, anticipating that additional series with both shorter and longer wavelengths probably existed, believed the designations K, L, M, etc. were preferable.
[66] H.G.J. Moseley, Phil Mag. **26** 1024 (1913); **27** 703 (1914).

The Bohr Hydrogen Atom

Since the known stability of atomic systems could not be reconciled with classical principles of mechanics and electrodynamics, Niels Bohr, in 1913, reasoned that classical physics laws must be wrong when applied to the motion of the electron in Rutherford's model of the atom.[67] Bohr coupled Rutherford's model with quantum theory to produce the first quantum theory of atomic structure.

A body that spins about its own axis or revolves in an orbit about a central point possesses angular momentum. Bohr assumed that the electron's angular momentum was restricted to certain values (he quantized the angular momentum). Each of the values, which was described by a principal quantum number, n, would specify a particular circular orbit. An atomic system, whose electrons were in given orbits, would not emit electromagnetic radiation even though the particles were accelerating. The whole atom was said to be in a stationary state. This assumption is contrary to classical electrodynamics as mentioned earlier. In Bohr's theory, electromagnetic radiation would be emitted or absorbed only when an electron changed from one allowed orbit to another allowed orbit. The energy difference between the two states would be emitted or absorbed in the form of a single quantum of radiant energy. This would produce the distinct spectral lines, described earlier, of a definite frequency, ν, related to the energy E by the relation $E = h\nu$ (already postulated by Planck and Einstein).

In Bohr's theory[68] for the hydrogen atom, the radii of the electron's circular orbits are given by

$$r = \frac{n^2\hbar^2}{m_r e^2}, \quad n = 1, 2, 3, \ldots$$

where m_r, the reduced mass, is given by

$$m_r = \frac{m_e M}{m_e + M}$$

m_e is the electron mass, M is the mass of the hydrogen nucleus, e is the charge on the electron, and $\hbar = h/2\pi$. The corresponding total energy of the electron is given by

$$E = -\frac{m_r e^4}{2\hbar^2 n^2}, \quad n = 1, 2, 3, \ldots$$

If an electron were to move from an orbit specified by the quantum number n_i to a different orbit specified by the quantum number n_f, the difference in energy of the electron would be

$$E_i - E_f = \frac{m_r e^4}{2\hbar^2}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

The frequency of the light emitted (or absorbed) during this transition would be given by

$$\nu = \frac{c}{\lambda} = \frac{(E_i - E_f)}{h} = \frac{m_r e^4}{4\pi\hbar^3}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

The wave number, k, (the reciprocal of the wavelength) is given by

$$k = \frac{1}{\lambda} = \frac{m_r e^4}{(4\pi\hbar^3)c}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

[67] N. Bohr, Phil. Mag. **26** 1 (1913); **26** 476 (1913).
[68] More complete discussions of the Bohr theory of the hydrogen atom can be found in many nuclear physics textbooks, e.g. R.G. Sachs, *Nuclear Theory*, (Addison-Wesley Publishing Company, Cambridge, Mass., 1953).

Period	Series	Group I a	Group I b	Group II a	Group II b	Group III a	Group III b	Group IV a	Group IV b	Group V a	Group V b	Group VI a	Group VI b	Group VII a	Group VII b	Group VIII a	Group O b
1	1	H 1															He 2
2	2	Li 3		Be 4			B 5		C 6		N 7		O 8		F 9		Ne 10
3	3	Na 11		Mg 12			Al 13		Si 14		P 15		S 16		Cl 17		Ar 18
4	4	K 19		Ca 20		Sc 21		Ti 22		V 23		Cr 24		Mn 25		Fe 26 Co 27 Ni 28	
	5		Cu 29		Zn 30		Ga 31		Ge 32		As 33		Se 34		Br 35		Kr 36
5	6	Rb 37		Sr 38		Y 39		Zr 40		Cb 41		Mo 42		43		Ru 44 Rh 45 Pd 46	
	7		Ag 47		Cd 48		In 49		Sn 50		Sb 51		Te 32		I 53		Xe 54
6	8	Cs 55		Ba 56		rare earths		72		Ta 73		W 74		75		Os 76 Ir 77 Pt 78	
	9		Au 79		Hg 80		Tl 81		Pb 82		Bi 83		Po 84		85		Rn 86
7	10	87		Ra 88		Ac 89		Th 90		91		U 92					

rare earths	La 57	Ce 58	Pr 59	Nd 60	61	Sm 62	Eu 63	Gd 64
	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70	Lu 71	

Figure 4. The Short Form of the Periodic Table, ca. 1913. *In the United States, the name columbium, Cb, was used for niobium, Nb (Z = 41), until 1949. In addition, by 1913, several claims as to the discovery of the element with Z = 43 had been made, but all were either proven false or never substantiated. (See Reference 13 for details.)*

The form of this equation is identical to that given by Rydberg (see Optical Spectroscopy, page 6) for the Balmer series of lines in the hydrogen optical spectrum if $n_f = 2$ and

$$R_H = \frac{m_r e^4}{4\pi c \hbar^3}$$

The value of R_H calculated from the above equation was found to be in good agreement with the value experimentally determined by Rydberg.[69]

Since other values of n_f are allowed in the above equation, the Bohr model predicted additional series in the optical spectrum of hydrogen. In succeeding years, four other series were observed experimentally: the Lyman[70] series in the ultraviolet region, corresponding to $n_f = 1$ and the Paschen[71], Brackett[72] and Pfund[73] series in the infrared region, corresponding to $n_f = 3$, 4, and 5, respectively. The Bohr model succeeded remarkably well in explaining the spectroscopic properties of the hydrogen atom.

Extension of the Bohr Theory

As higher resolution spectroscopic studies of hydrogen were made, it was found that spectral lines, originally believed to be single lines, actually consisted of a number of closely spaced lines. To explain this *fine structure*, the quantum theory was further refined in 1916 when Arnold Sommerfeld introduced[74] an azimuthal quantum number, ℓ, where $\ell \leq n - 1$. With this modification discrete elliptical orbits, in addition to the circular orbits, are allowed for electrons. This change permitted the Bohr model to account for detailed structure in the pattern of radiation emitted by hydrogen and other atoms.

To explain the change in the emitted radiation pattern when an atom is exposed to a magnetic field, a magnetic quantum number, m_ℓ, with permitted values from $-\ell$ to $+\ell$, was added. This quantum number designates different projections of the possible circular or elliptical orbits along the direction of the magnetic field in space.

Finally, to account for the close grouping of two or more spectral lines, Samuel Goudsmit and George Uhlenbeck postulated that the electron has an intrinsic angular momentum, called spin.[75] In a magnetic field, the spin axis can have two directions relative to the field, corresponding to the two possible values for the spin quantum number $m_s = \pm 1/2$.

All particles or aggregations of particles have a net spin which is either an integer (including zero) or half-integer. Particles with half-integer spin, including electrons and protons, obey the Pauli exclusion principle.[76] In general terms, this principle states that, in a given system, two identical particles cannot have the same set of quantum numbers. (The behavior of large numbers of particles with half-integer spin is described by Fermi-Dirac statistics, and the particles are collectively called fermions.[77] Particles with integral spin, such as photons, do not obey the Pauli principle. Instead of Fermi-Dirac statistics, they obey Bose-Einstein statistics and are collectively called bosons. At high temperature or low densities, both the Fermi-Dirac and Bose-Einstein behaviors converge toward classical, or Boltzmann statistics.)

For multi-electron atoms, the distribution of electrons among the allowed orbits was first studied by N. Bohr[78] and E. Stoner.[79] The orbits are determined by a set of the four quantum numbers,

[69] The Rydberg constant, $R_\infty = m_e e^4/(4\pi\varepsilon_0^2 h^3)$, is obtained from R_H as $M/m_{e-\infty}$.
[70] T. Lyman, Nature **93** 241 (1914).
[71] F. Paschen, Ann. d. Phys. **27** 565 (1908).
[72] F.S. Brackett, Nature **109** 209 (1922).
[73] A.H. Pfund, J. Opt. Soc. Am. **9** 193 (1924).
[74] A. Sommerfeld, Ann. d. Phys. **51** 1 (1916).
[75] G.E. Uhlenbeck and S. Goudsmit, Naturwiss. **13** 953 (1925).
[76] W. Pauli, Z. Physik **31** 765 (1925).
[77] The terms "fermion" and "boson" are contractions of the terms "Fermi-Dirac particle" and "Bose-Einstein particle". See R.D. Evans, *The Atomic Nucleus*, (McGraw Hill Book Company, New York, 1955), p. 180.
[78] N. Bohr, Zeit. f. Phys. **9** 1 (1922).
[79] E.C. Stoner, Phil. Mag. **48** 719 (1924).

n, ℓ, m_ℓ, and m_s. All electrons with principal quantum number n = 1 are in the innermost orbit, called the K-shell.[65] Because electrons obey the Pauli principle, only two electrons can occupy this shell. Both will have $\ell = 0$ and $m_\ell = 0$, but m_s can be ±1/2. For historic reasons, the integral values of the quantum number ℓ are often represented by letter symbols, i.e. s, p, d, f... corresponding to $\ell = 0, 1, 2, 3\ldots$ respectively.[80] The innermost shell is designated the 1s shell (or subshell) corresponding to n = 1 and $\ell = 0$. For n = 2, the L-shell, two subshells are possible, the 2s and 2p. The designation of the shells and subshells, as well as the maximum number of electrons in each is shown in Table 1.

Table 1. Electron Subshells for n = 1 - 4.

n	1	2	3	4
shell	K	L	M	N
ℓ	0	0, 1	0, 1, 2	0, 1, 2, 3
subshell	1s	2s, 2p	3s, 3p, 3d	4s, 4p, 4d, 4f
electrons/ subshell	2	2, 6	2, 6, 10	2, 6, 10, 14
total electrons	2	8	18	32

Electron Configurations

The structure of the atom predicted by the Bohr model provided a satisfactory understanding of many atomic phenomena. A neutral atom of an element has a characteristic number, Z, of net positive electronic charges in the nucleus with an equal number of electrons in the electron shells to balance the charge. Ions, the species responsible for electrical conduction in Faraday's experiments[5], are simply atoms that have more or fewer electrons in their electron shells than positive charges in their nuclei. The number of protons in the nucleus determines the elemental identity of the atom, while the difference between the number of electrons in the electron shells and the nuclear charge determines whether the atom behaves as a neutral atom, a cation, or an anion.

When enough electrons are present to fill an electron shell, the atom is in a stable configuration and does not easily undergo chemical reactions. The inert gases have this electron configuration.

The similarities in chemical properties of the elements in the same column of Mendeleev's periodic table were due to the similarities in the configurations of the outer electron shells for those materials. Hydrogen, lithium, and other elements in the first column of the periodic table have only one electron in the outermost shell; beryllium, magnesium, calcium and the other elements in the second column have two electrons in the outermost shell.

The similarity in the chemical properties of the rare earth elements was explained by the similarities in the structure of the outer electron shells for these materials. The additional orbital electrons in the elements from lanthanum to hafnium fill an inner electron shell (the 4f shell), while the outer shell structure of these elements stays almost the same.

The groups of X-ray lines, the K-series and the L-series, studied by Barkla[54] and Moseley[68], can now be explained in terms of transitions within the shell structure of the electrons. If an electron is removed from one of the inner shells of an atom, an electron in some outer shell can fill the vacancy. The final state of the atom will be of lower energy and the atom would emit the energy difference in the form of an X-ray. The K-series of X-rays is produced by vacancies in the K-shell being filled by electrons originally in the L- or M-shell. The K_α series is due to transitions from the L subshells to the K-shell, while the K_β series arises from transitions from the M subshells to the K-shell. Similarly, the L_α and L_β series are produced by transitions from the M subshells to the L subshells. Because of the large energy difference between these shells (on the order of keV), the radiated X-rays are very penetrating, having energies that are ~1000 times greater than visible light.

Wave-Particle Duality

The wave nature of light had been demonstrated in experiments involving reflection, refraction and diffraction. As mentioned earlier, Freidrich, Knipping and von Laue demonstrated the wave nature of X-rays by showing that X-rays are diffracted by crystals.[32] Other studies, mentioned above, seemed to indicate that light also has a particle nature. In Planck's treatment of blackbody radiation, the concept of discrete amounts of energy (the energy quantum) was first introduced. Einstein, then, applied this concept to light in explaining the photoelectric effect.

The quantum theory states that electromagnetic radiation must be emitted or absorbed in integral multiples of these quanta. The quantization of its energy gives electromagnetic radiation some of the properties of particles and the term "photon" is used to refer to such particles of radiation.

In analogy with the wave-particle properties of light, Louis de Broglie hypothesized in 1924 that particles might have wave properties as well.[81] Like photons, the momentum and wavelength of particles would be related by $p = h/\lambda$. He further showed that the allowed orbits of the electrons in the Bohr model of the hydrogen atom could be defined by the condition that the circumference of the electron's orbit contains an integral number of the electron's wavelengths, i.e. standing waves.[82]

In 1927, C. Davisson and G. Germer showed that electrons could be diffracted by crystal lattices.[83] Other experiments also showed that when electrons passed through thin films of metal foils, the resulting diffraction patterns were similar to those obtained with X-rays.[84] These experiments confirmed de Broglie's postulate.

Quantum Mechanics

As pointed out earlier, several phenomena could not be explained using the laws of classical physics and ad hoc assumptions were required, for example, Bohr's assumption that the orbital angular momentum of the electron was restricted to certain discrete values. A new departure was provided by the development of quantum mechanics almost simultaneously by W. Heisenberg[85] and by E. Schrödinger.[86] Starting with de Broglie's postulate, these theories describe the motion of the de Broglie waves of a particle under the influence of an externally applied potential. The stationary states of the Bohr atom are, then, conceived as proper solutions of a fundamental wave equation.

The Heisenberg Uncertainty Principle

Another result of the de Broglie hypothesis was the Heisenberg uncertainty principle.[87] Heisenberg stated that it is impossible to simultaneously determine both the position and momentum of a particle exactly. This limitation is expressed mathematically as

$$\Delta x \Delta p_x \geq \hbar .$$

Decreasing the uncertainty in the position, Δx, of the particle results in increased uncertainty in the momentum, Δp_x, of the particle. One of the causes of the uncertainty relationship is that

[80] R.B. Leighton, *Principles of Modern Physics*, (McGraw-Hill Book Company, Inc., New York, 1959), p. 204 and G. Herzberg, *Atomic Spectra and Atomic Structure*, (Dover Publications, New York, 1944), p. 86.

[81] L. de Broglie, Phil. Mag. **47** 446 (1924).

[82] see R.B. Leighton, op. cit., p. 81.

[83] C. Davisson and L.H. Germer, Phys. Rev. **30** 705 (1927).

[84] G.P. Thomson, Proc. Roy. Soc. (London), **A117** 600 (1928); **A119** 663 (1928).

[85] W. Heisenberg, Z. Physik **33** 879 (1925).

[86] E. Schrödinger, Ann. Phys. **79** 361 (1926); **79** 489 (1926); **81** 109 (1926).

[87] W. Heisenberg, Z. Physik **43** 172 (1927).

in measuring a property of a particle, e.g. the position, the measurement process, itself, disturbs the particle whose properties are being measured.

Similar relationships apply to certain other pairs of quantities, for example, energy and time. If the energy, E, of a particle is measured in a time interval, Δt, there will be a resulting uncertainty in the energy, ΔE, where

$$\Delta E \Delta t \geq \hbar \, .$$

Discovery of the Neutron

Through the early 1900s, the structure of the nucleus was not well understood. The lightest positively charged particle was, presumably, the hydrogen nucleus or proton.[24] Though scientists searched for a lighter particle, none was found. A proton-electron model had been assumed for the nucleus.[88] Nuclei of atoms heavier than hydrogen might be composed of protons and electrons, the protons supplying the bulk of the mass, while the electrons balance the electronic charge. The mass, A, of a given nucleus, expressed in units of the proton mass, would be approximately 2Z, where Z is the position of the element on the periodic table. A nucleus might then be composed of A protons and (A-Z) electrons. A neutral oxygen atom, for example, with Z = 8, would have a nucleus consisting of 16 protons and 8 electrons, and would have eight orbital electrons.

This view of the structure of the nucleus was altered with the discovery of the neutron. In 1930, W. Bothe and H. Becker found that a highly penetrating radiation was emitted when beryllium, boron or lithium were bombarded by alpha particles from polonium.[89] The radiation was not affected by magnetic fields and was not ionizing radiation. Initially, it was believed that this radiation might be high-energy gamma-rays. In subsequent experiments carried out by Irene Curie and Frederic Joliot in 1932, protons were found to be produced when this radiation struck hydrogen-containing substances such as paraffin.[90] James Chadwick demonstrated the evidence from these experiments was compatible with the assumption that the radiation consisted of an uncharged particle having a mass approximately equal to that of the proton.[91] The particle was called the neutron.

Soon after the discovery of the neutron, Heisenberg suggested a neutron-proton model of the nucleus.[92] The nucleus consists of Z protons and N neutrons. Protons and neutrons are both referred to as nucleons. The total number of nucleons in the nucleus is called the mass number, A = N + Z. An isotope is one of a group of two or more nuclides having the same number of protons or the same atomic number, but different numbers of neutrons. For example, natural hydrogen consists almost entirely of atoms having nuclei consisting of a single proton. However, a small amount (0.0115%) of deuterium (heavy hydrogen) is present in nature; the deuterium atom has a nucleus consisting of one proton and one neutron. The hydrogen and deuterium atoms both have a single orbital electron. In general the situation becomes more complex as the heavier elements are encountered. Natural tin, which has an atomic number of Z = 50, consists of 10 stable isotopes of masses 112, 114, 115, 116, 117, 118, 119, 120, 122, and 124. Although the nuclei of all these isotopes have 50 protons, each contains a different number of neutrons (ranging from 62 to 74). The neutral atoms of these isotopes all have 50 orbital electrons.

The term "nuclide" refers to a species of atom characterized by the number of neutrons and protons contained in the nucleus.[93] Symbolically, a nuclide of element "X" may be represented by the elemental symbol with a subscript specifying Z and a superscript specifying A:

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

For example, the isotope of tin with A = 115 is represented as

$$^{115}_{50}Sn$$

As mentioned above, in the study of the naturally-occurring radioactive series, more than 40 radioactive "species" were isolated. Soddy had pointed out that several of these materials were chemically identical and, with the new understanding of isotopes, the reasons were clear. The materials ionium, radiothorium and thorium were all isotopes of thorium, ^{230}Th, ^{228}Th, and ^{232}Th. RaA, AcA, and ThA were all isotopes of polonium; RaB, AcB, and ThB were all isotopes of lead. The three naturally-occurring radioactive decay series[94] are discussed in more detail on page 24.

The Positron

Paul Adrien Maurice Dirac first postulated the existence of an antiparticle, the positron, in 1931.[95] The positively charged positrons would have the same mass (511 keV)[96] and magnitude of charge as the electron and would react with electrons in matter by annihilation to produce gamma radiation. In an annihilation reaction, two 511 keV gamma-rays, emitted in opposite directions, are produced, conserving momentum and energy. In studies of cloud chamber photographs of cosmic rays, C.D. Anderson established the existence of these particles in 1932.[97] This was the first indication that the universe is made up of both matter and antimatter.

Artificial Radioactivity

In 1932, after learning of the discovery of the positron, I. Curie and F. Joliot re-examined experiments they had carried out earlier on the bombardment of certain elements with alpha particles from polonium sources.[98] For some light elements, notably beryllium and aluminum, not only were neutrons emitted, but positrons were also observed. Furthermore, when the alpha source was removed, the emission of neutrons ceased but the positron emission continued, decaying with a characteristic half-life. With aluminum as the target, the half-life of the positron emission was about three minutes. For aluminum, Curie and Joliot hypothesized that the absorption of an alpha particle, ^{4}He, by aluminum, ^{27}Al, and subsequent emission of a neutron, ^{1}n, would lead to the formation of an isotope of phosphorus, ^{30}P. The phosphorus might be the source of the positron emission, decaying to ^{30}Si. While the detection of the small amounts of silicon produced in this reaction was not possible by chemical means, other chemical separations to isolate the phosphorus were devised. The positron emission followed the phosphorus in these separations, indicating the hypothesis was correct.

The isotope ^{30}P does not exist in nature. The work of Curie and Joliot showed that new isotopes could be produced through nuclear reactions and they introduced the term "artificial radioactivity" to indicate the manner in which these new isotopes were produced.

Nuclear Reactions

The initial reaction studied by Curie and Joliot can be

[88] W.D. Harkins and E.D. Wilson, J. Am. Chem. Soc. **37** 1367 (1915); **37** 1383 (1915); **37** 1396 (1915).
[89] W. Bothe and H. Becker, Zeits. f. Phys. **66** 289 (1930).
[90] I. Curie and F. Joliot, Compt. Rend. **194** 273 (1932).
[91] J. Chadwick, Proc. Roy. Soc. (London) **A136** 692 (1932).
[92] W. Heisenberg, Zeit. Phys. **78** 150 (1932).
[93] The word "nuclide" was proposed by Truman P. Kohman. See T.P. Kohman, Am. J. Phys. **15** 356 (1947).

[94] The symbols for the historic names of the nuclides belonging to each of the three naturally-occurring decay series are given in the squares for each of these nuclides on the Chart of the Nuclides.
[95] P.A.M. Dirac, Proc. Roy. Soc. **A133** 60 (1931).
[96] The electron volt (eV) is defined as the energy necessary to raise one electron through a potential difference of 1 volt. The kiloelectron volt (keV) equals 1000 eV and the megaelectron volt (MeV) equals 10^6 electron volts.
[97] C.D. Anderson, Phys. Rev. **43** 491 (1933).
[98] I. Curie and F. Joliot, Compt. Rend. **194** 708 (1932).

represented by the following equation:

$$^{27}_{13}Al+^{4}_{2}He\rightarrow^{1}_{0}n+^{30}_{15}P$$

The symbols on the left represent the reacting nuclides; those on the right are the products of the reaction. Note that the superscripts on one side of the arrow balance those on the other side. The same is true for the subscripts. The balance represents the conservation of the number of protons and neutrons separately. In addition to charge and mass number, all nuclear reactions conserve momentum, angular momentum, mass-energy, and spin. A more abbreviated notation is frequently used:

$$^{27}Al(\alpha,n)^{30}P$$

In 1935, a summary of the known nuclides and the reactions used to study the nuclides was published by Giorgio Fea.[99] Included in the work was a "proton-neutron chart," showing the known isotopes of each element, whether the nuclide was stable or radioactive, and, if stable, an indication of the relative abundance of the isotope. A portion of Fea's chart is shown in Figure 5. This was one of the earliest published, "Chart of the Nuclides".

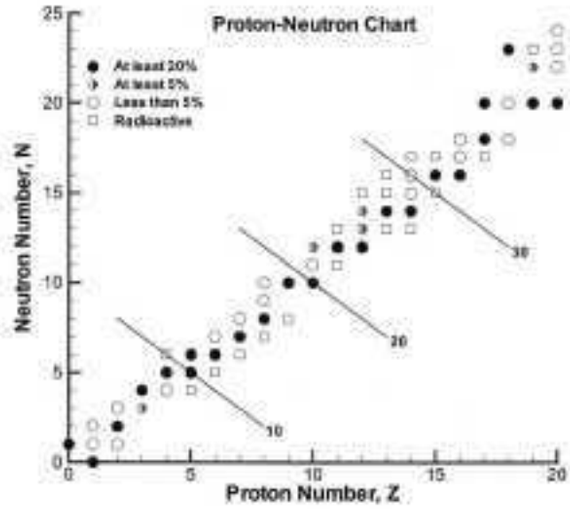

Figure 5. A Portion of the Proton-Neutron Chart Published by G. Fea in 1935 (Ref. 99). *Radioactive nuclides are indicated by squares. The relative abundance of the stable nuclides is indicated by the shading of the circles as shown in the legend.*

Beta Decay and the Neutrino

The term "beta radiation" given to the radiation observed from the uranium and thorium compounds by Rutherford includes two types of decay: β^-, the emission of an electron from a nucleus, and β^+, the emission of a positron from a nucleus. These forms of decay differ from alpha, gamma, and X-ray emission in that β^- and β^+ particles have continuous spectra of energies ranging from near zero up to an end-point value that is characteristic of the decaying nuclide. To account for the continuous spectrum Wolfgang Pauli suggested[100] in 1930 that a new type of particle, the neutrino (little neutral one)[101] or its antiparticle, the antineutrino, might be emitted in beta decay. These are uncharged particles with either extremely small or zero rest mass and have extremely low probabilities of interacting with matter. The sum of the beta particle energy and the neutrino or antineutrino energy, rather than the beta particle energy alone, would have discrete values.

The production of antineutrinos from the beta decay of the neutron was confirmed in 1956. The detection of the products of the reaction of antineutrinos (produced from the decay of neutrons in a high flux reactor) and protons by Reines and Cowan gave unmistakable evidence for the existence of the antineutrino.[102]

For β^- decay, the process is represented by

$$^{A}_{Z}X\rightarrow^{A}_{Z+1}Y+^{0}_{-1}e+^{0}_{0}\overline{\nu}$$

or

$$^{A}_{Z}X\rightarrow^{A}_{Z+1}Y+\beta^- +\overline{\nu}$$

where $\overline{\nu}$ is an antineutrino. Likewise, for β^+ decay

$$^{A}_{Z}X\rightarrow^{A}_{Z-1}Y+^{0}_{1}e+^{0}_{0}\nu$$

or

$$^{A}_{Z}X\rightarrow^{A}_{Z-1}Y+\beta^+ +\nu$$

where ν is a neutrino.

An alternative to β^+ decay, electron capture, ε, was proposed by Yukawa and Sakata in 1936.[103] Here an orbital electron from the K or L shell is captured by the nucleus and is accompanied by the emission of a neutrino.

$$^{A}_{Z}X+^{0}_{-1}e\rightarrow^{A}_{Z-1}Y+^{0}_{0}\nu$$

As in β^+ decay, the atomic number, Z, of the original nucleus is reduced by one, while the mass number, A, is unchanged. Alvarez observed this type of decay in 1938.[104]

Production of New Elements Using Nuclear Reactions

The realization that new nuclides could be produced through reactions with alpha particles led to the development of particle accelerators capable of producing beams of high-energy charged particles. Among these machines were the Cockcroft-Walton accelerator,[105] the Van de Graaff accelerator,[106] the cyclotron,[107] the betatron[108] and the linear accelerator[109] (linac). Reactions induced by protons, deuterons, neutrons (produced using charged-particle induced reactions) and photons were shown to lead to other radioactive products, each characterized by a half-life and certain decay products.

Nuclear reactions were used to produce the elements still missing from the periodic table below Z = 92. In 1937, the element technetium (from the Greek word "technetos" meaning artificial) was produced by bombarding a molybdenum target with deuterons.[110] This was the first element to be produced artificially. Similarly, the element promethium was produced in 1941 by bombarding a neodymium target with deuterons.[111] No naturally-occurring (stable) isotopes of either technetium or promethium exist.

Fission

In attempts to produce transuranium elements, i.e. elements

[99] Georgio Fea, Nuovo Cimento **2** 368 (1935).
[100] W. Pauli, unpublished letter, Dec. 4, 1930. An English translation is given in J. Steinberger, Science **259** 1872 (1993).
[101] The name "neutrino" was suggested by Enrico Fermi. See Emilio Segre, *Enrico Fermi, Physicist*, (University of Chicago Press, Chicago 1970), p. 70.
[102] F. Reines and C.L. Cowan, Jr., Nature **178** 446 (1956); Phys. Rev. **113** 273 (1959).
[103] H. Yukawa and S. Sakata, Proc. Phys. Math. Soc. Japan. **18** 128 (1936).
[104] L. Alvarez, Phys. Rev. **53** 213 (1938); **54** 486 (1938).
[105] J.D. Cockcroft and E.T.S. Walton, Proc. Royal. Soc. **A136** 619 (1932).
[106] R.J. Van de Graaff, Phys. Rev. **38** 1919 (1931).
[107] E.O. Lawrence and N.E. Edlefsen, Science **62** 376 (1930); E.O. Lawrence and M.S. Livingston, Phys. Rev. **40** 19 (1932).
[108] D.W. Kerst, Phys. Rev. **60** 47 (1941).
[109] R. Wideröe, Arch. Elektrotech. **21** 387 (1928).
[110] C. Perrier and E. Segre, Nature **140** 193 (1937).
[111] J.A. Marinsky, L.E. Glendenin and C.D. Coryell, J. Am. Chem. Soc. **69** 2781 (1947).

beyond uranium in the periodic table, scientists bombarded uranium targets with slow neutrons, believing the capture of a neutron by the uranium nucleus, followed by beta decay, might produce isotopes of elements with Z > 92. They expected the atomic number of the resulting nuclides to be only slightly different from that of uranium. The experiments were complicated by the difficulty in chemically separating the trace amounts of reaction products. Furthermore, choosing the "proper" chemical techniques for separation was difficult since the chemistry of the transuranium elements was open to speculation. Correlating the properties of the newly produced radioisotopes with those predicted for the transuranium elements proved extremely difficult. For several years, scientists were misled by the a priori assumption that the neutron bombardment of uranium would produce only transuranic elements.

In 1939, Hahn and Strassmann showed that the products of the neutron bombardment of uranium included barium, lanthanum and cerium, with Z = 56, 57 and 58, respectively.[112] The presence of elements in the medium mass region was totally unexpected. Other medium mass materials—molybdenum, rubidium, antimony, and iodine—were detected in further chemical analyses. The presence of elements of intermediate mass indicated the uranium nucleus had split into fragments. L. Meitner and O. Frisch called this new reaction fission.[113] They also showed that the energy associated with the fission of the ^{235}U nucleus was on the order of 200 MeV.

The emission of neutrons during fission was soon reported by several groups of researchers.[114] The possibilities of a chain reaction with the liberation of a tremendous amount of energy became apparent. With the obvious military possibilities of nuclear energy, a secret program for the development of a fission bomb was begun in 1939 in the United States.[115]

Production of Elements with Z = 93 and 94

The elements with Z greater than 92 were produced by other nuclear reactions. In 1940, E.M. McMillan and P.H. Abelson identified[116] an isotope of element 93, ^{239}Np, produced by the capture of a neutron by ^{238}U:

$$^{238}_{92}\mathrm{U} + ^{1}_{0}\mathrm{n} \rightarrow ^{239}_{92}\mathrm{U} + \gamma$$

and

$$^{239}_{92}\mathrm{U} \rightarrow ^{239}_{93}\mathrm{Np} + \beta^{-} + \bar{\nu}$$

The new nuclide, ^{239}Np, was itself radioactive, decaying by β^- emission with a reported half-life of 2.4 days. Since this new element was the first element beyond uranium, which was named after the planet Uranus, it was named neptunium, from Neptune, the first planet beyond Uranus.

Plutonium with Z = 94 was produced by bombarding ^{238}U with deuterons from the 60-inch cyclotron of the University of California at Berkeley by G.T. Seaborg, E.M. McMillan, J.W. Kennedy, and A.C. Wahl in 1941.[117] The new element was named after the planet Pluto, following the pattern used for neptunium.

Modifications to the Periodic Table

The newly discovered elements, neptunium (Z = 93) and plutonium (Z = 94), were initially placed in the seventh row of the periodic table, following uranium (Z = 92). It was assumed that elements 95 and 96 would have chemical properties similar to those of neptunium and plutonium. Attempts were made[118] to produce these elements by bombarding ^{239}Pu with neutrons and deuterons and chemically isolating the new elements. These experiments failed.

It was suggested that the elements following actinium on the periodic table should be placed in a second series, similar to the series of rare earths (which follow lanthanum).[119] If this were the case, the chemical properties of elements 95 and 96 would be similar to actinium (Z = 89) and to the corresponding elements in the rare earths, europium (Z = 63) and gadolinium (Z = 64). With this assumption, new experiments[118] using different methods for chemically isolating the products were carried out[120] and succeeded[121] in producing isotopes of elements 95 (^{241}Am), in 1944, and 96 (^{242}Cm), in 1945.

In this form of the periodic table, the rare earth series, starting with cerium (Z = 58) and ending with lutetium (Z = 71), was called the Lanthanide Series. The newer series[122] was located below the lanthanides and was called the Actinide Series. This new series began with thorium (Z = 90) and would, ultimately, end with lawrencium, Z = 103.

Recent Work

Since 1945, isotopes of eighteen new elements have been produced. This work is summarized in "An Island of Stability for Superheavy Elements," beginning on page 31. Names have been given to fourteen of the eighteen new elements. The known elements are listed in order of increasing atomic number in the table on the inside front cover of this booklet and alphabetically by name on page 48. In both lists, the atomic masses[123] (in atomic mass units, u) and the number of isotopes[124] are given for the stable and naturally-occurring radioactive elements. In the tables, the number in parentheses following the mass is the uncertainty in the last significant figure of the mass.

The proton-neutron chart published by Georgio Fea in 1935 (Figure 5, page 13) included information on about 320 nuclides. Today, more than 3100 nuclides are known. The Chart of the Nuclides summarizes the known properties of these nuclides. A description of the type of information included on the Chart begins on page 17.

The Periodic Table of the Elements

The Periodic Table of the Elements is shown on the inside back cover of this booklet. The elements are arranged in the table, first, in order of increasing atomic number, the number of protons in the nucleus or the number of extranuclear electrons in the uncharged atom, and, then, in horizontal rows, called periods, and vertical columns, called groups. Members of each group have similar physical and chemical properties: Group 1 elements are called the alkali metals, Group 2 the alkaline earths, Group 17 the halogens, Group 18 the noble gases, etc. In this form of the

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[112] O. Hahn and F. Strassmann, Naturw., **27** 11 (1939).

[113] L. Meitner and O.R. Frisch, Nature **143** 239 (1939).

[114] H. von Halban, F. Joliot and L. Kowarski, J. Phys. Rad. **10** 428 (1939); H.L. Anderson, E. Fermi, and H.B. Hanstein, Phys. Rev. **55** 797 (1939); L. Szilard and W. Zinn, Phys. Rev. **55** 799 (1939).

[115] See Richard Rhodes, *The Making of the Atomic Bomb*, (Simon and Schuster, Inc., New York, 1986).

[116] E.M. McMillan and P.H. Abelson, Phys. Rev. **57** 1185 (1940).

[117] G.T. Seaborg, A.C. Wahl and J.W. Kennedy, Phys. Rev. **69** 367 (1946). (This work was submitted for publication in 1941, but was voluntarily withheld from publication until the end of World War II because of the nature of its contents.)

[118] G.T. Seaborg and W.D. Loveland; *The Elements Beyond Uranium*, (John Wiley & Sons, Inc. New York, 1990).

[119] G.T. Seaborg, Chem. and Eng. News **23** 2190 (1945).

[120] The chemical separation and identification of these elements were extremely difficult and went on for nearly a year. During this time, one member of Seaborg's research group referred to these (as yet unnamed) elements as "pandemonium" and "delirium". See M.E. Weeks and H.M. Leicester, *Discovery of the Elements*, (Mack Printing Company, Easton, Pa., 1968), p. 847.

[121] The first public announcement of the discovery of americium and curium was made by Seaborg on the Quiz Kids radio program, November 11, 1945. See G.T. Seaborg, J. Chem. Ed **45** 278 (1968).

[122] See Reference 121 for a discussion of changes to the periodic table during this time.

[123] As mentioned earlier, Dalton obtained masses relative to the hydrogen atom. In 1819, Berzelius published a table of masses relative to the mass of oxygen, which was taken to be 100. This system was used until the 1860's, when chemists again adopted the hydrogen scale. In 1906, this was converted back to the oxygen scale and the atomic mass unit was taken to be 1/16 of the atomic mass of oxygen. After the discovery of the oxygen isotopes, two scales were used—a chemical scale based on 1/16 of the average mass of the natural oxygen atom and a physical scale, using a mass unit equal to 1/16 of the mass of the ^{16}O atom. In 1961, a single scale based on ^{12}C, whose mass was taken to be 12.000000, was chosen. Elemental masses given in the present work are taken from T.B. Coplen, Pure Appl. Chem. **73** 667 (2001) and Chem. Int. **23** 179 (2001).

[124] Eighty-four elements have stable or long-lived naturally-occurring isotopes. There are a total of 288 stable or long-lived naturally-occurring isotopes for which isotopic abundances are given on the Chart of the Nuclides. See "Isotopic Abundances" on page 19.

periodic table, two separate periods of elements, the Lanthanides and the Actinides, are shown outside the main table.

Within each square, the atomic number, the name and symbol, the relative atomic mass, and several physical and chemical properties are given for the element. The color of the name and symbol indicates the form of the element at standard temperature and pressure[125] (STP): gas (green), liquid (blue), or solid (black). Elements that do not exist in nature (artificially prepared) are shown in red. The atomic number is given in the upper left corner of the square and is the number of protons in the nucleus of the atom or the number of orbital electrons in the uncharged atom. The relative atomic masses of the naturally-occurring elements, in atomic mass units (u), are given below the symbols. The atomic mass unit is 1/12 of the mass of the ^{12}C atom in its atomic and nuclear ground state.

In the upper right corner of the square, the common oxidation states of the element are given with the most common in red.[126] For an element in a combined state (a compound), the oxidation state is the number of electrons that must be added to or subtracted from the ion to convert it to the elemental form. For example, in the compound $NiCl_2$, the oxidation state of Ni is +2, while that of chlorine is –1. While some elements have only one oxidation state (see the Group 1 and 2 elements in the Periodic Table) others can take on several values (see Groups 7, 8, and 9, for example).

Below the oxidation states, the electronegativities (from Lide)[126] of the elements are given. These are parameters which describe, on a relative basis, the tendencies of an atom in a molecule to attract bonding electrons and are discussed in the Chemical Bonding section, beginning on page 15.

Along the lower left side of each square, the melting and boiling points, the specific gravity and the thermal conductivity are given for each element.

- Melting and Boiling Points. The melting and boiling points (from Lide)[126] are given in degrees Celsius and are at atmospheric pressure. For some materials, e.g. carbon and arsenic, the sublimation temperature (denoted by "subl") is given instead of the boiling point. The melting and boiling points of the elements having $18 \le Z \le 90$ are shown in Figure 6. Three periods are shown on the graph: $19 \le Z \le 36$, $37 \le Z \le 54$, and $55 \le Z \le 86$. The last period includes the Lanthanides ($58 \le Z \le 71$). Similarities in the trends of the melting and boiling points for the three periods should be noted.

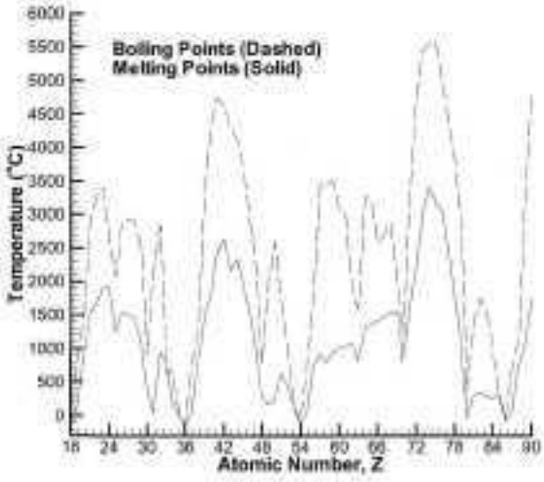


Figure 6. Melting and Boiling Points for the Elements Between Z = 18 and Z = 90.

- Specific Gravity (For Gases, Density). For elements that are liquids or solids at standard temperature and pressure (STP), the specific gravity is given. The specific gravity of a material is the ratio of the mass of that material to the mass of an equal volume of water. The specific gravities (from Lide)[126] given on the periodic table are at 20°C. For elements that are gases at STP, the density is given in grams per liter at 0°C.
- Thermal Conductivity. The thermal conductivity of a material is the rate of heat transfer by conduction through a unit thickness per unit area per unit difference of temperature and is given in units of watts per meter per degree Celsius.[127]

On the lower right side of each square, the atomic radius, the principle X-ray emission energies in keV, and the crustal abundance are given.

- Atomic Radius. The atomic radius (from Emsley)[127] is given in picometers (10^{-12} meters). Generally speaking, the atomic radius increases from top to bottom of a group and decreases from left to right over a period. Figure 7 is a plot of the atomic radius for elements with $18 \le Z \le 90$. Note the similarities in the shape of the curve going from period to period.

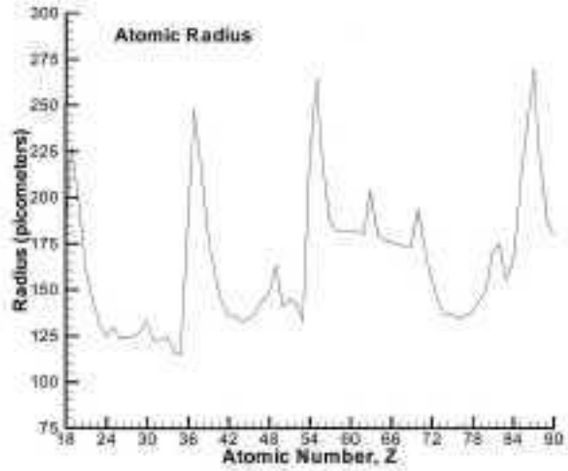


Figure 7. The Atomic Radius for the Elements Between Z = 18 and Z = 90. *The peaks occur at Z = 19, 37, 55, and 87, corresponding to elements in Group 1 of the Periodic Table.*

- Principal X-ray energies.[128] For elements with $Z \le 62$, energies of the K_α and K_β lines are given, while for $Z > 62$, the energies correspond to the L_α and L_β lines. The regular variation (observed by Moseley)[66] of the square root of the energy of these X-rays as a function of the atomic number is shown in Figure 8, page 16.
- Crustal Abundance. Geologically, the planet earth consists of three shells, the core, the mantle and the crust. Though the crust is only about 20 km thick and contains about 1% of the total mass of the earth, this region is vitally important since it provides the raw materials required by man. The crustal abundance (from Lide)[126] of each element is given in parts per million by mass.

Chemical Bonding

In the first half of the 20th century, G.N. Lewis,[129] Irving

[125] "Standard Temperature and Pressure" (STP) is defined as a temperature of 0°C (273.15 K) and a pressure of 1 atmosphere (101.325 kPa).

[126] David R. Lide, Editor-in Chief, *Handbook of Chemistry and Physics, 81st Edition 2000-2001*, (CRC Press, Boca Raton, 2000).

[127] J. Emsley, *The Elements*, (Clarendon Press, Oxford, 1998).

[128] G.C. Johnson and G.B. White, *X-ray Emission Wavelengths and keV Tables for Nondispersive Analysis*, ASTM Data Series, DS46 (1970).

[129] G.N. Lewis, J. Amer. Chem. Soc. 38 762 (1916).

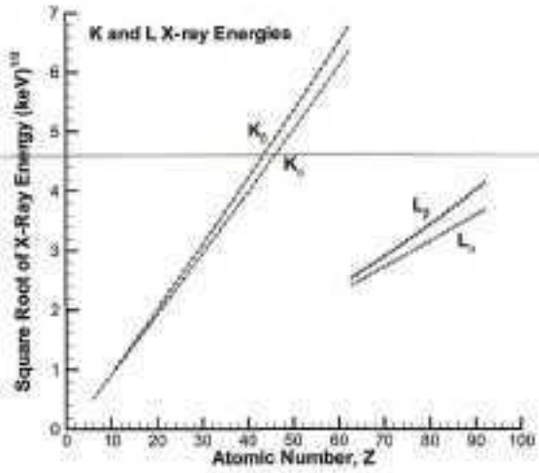


Figure 8. The K and L X-ray Energies Plotted as a Function of the Atomic Number. *The regularity of the square root of the X-ray energy as a function of the position of the element in the periodic table was first observed by Henry G.J. Moseley (see Studies of X-Ray Spectra, page 9).*

Langmuir,[130] Linus Pauling[131] and others explained much of what had been observed about the structure and stability of chemical substances in terms of atomic theory and wave mechanics. They recognized three idealized forms of strong primary bonding and several weaker secondary interactions. The three forms of primary bonding are called ionic, covalent, and metallic. The actual bonding in many compounds is best described as a combination of the primary forms.

Electrons are transferred in ionic bonding. If only one or two electrons are in the last unfilled shell of an element, it is relatively easy for the atom to lose these electrons to another atom whose last unfilled shell has one or two vacancies. The first of these atoms becomes a positively charged ion because of the loss of electrons, and the second becomes a negatively charged ion because of the gain of electrons. The electrostatic attraction between the ions binds them into a compound. The attraction is strongest for small ions with high charge. In idealized ionic bonding, the electron distribution around either type of ion is spherically symmetric. In ionic crystals, the relative sizes and numbers of the different types of ions largely determine their geometric arrangement. In an ionic crystal of formula A_nB_m, where $n \geq m$, the ions of element A are typically surrounded by four to eight ions of B in a lattice that resembles a three-dimensional checkerboard.

Atoms also bond to each other by sharing electrons, usually in pairs. Such covalent bonds are formed between atoms of the same element, or between pairs of elements that are not as far apart on the periodic table as the pairs that form ionic bonds. Atoms usually form just enough covalent bonds to fill their unfilled shells when each shared electron is counted as if it belonged to each atom sharing it. Covalent bonds are localized in space, in that the density of electrons is greater near a line between the two atoms than elsewhere, and pairs of covalent bonds to a given atom tend to have a characteristic angle between them. In typical covalent compounds, four or fewer nearest-neighbor atoms surround each atom. Atoms of different elements share electrons unequally. The tendency to attract shared electrons is called electronegativity. Although various definitions and scales of electronegativity have been proposed, they all agree that elements that most easily form negative ions have the highest electronegativity, while elements that most easily form positive ions have the lowest. Covalent bonds between atoms with different electronegativities are called polar covalent or partially ionic bonds. They gain additional strength through their partial ionic character.

Atoms in metals are bound by electrons that are spread out through the entire solid or liquid. Like ionic bonding, and unlike covalent bonding, metallic bonding has no preferred directions. Most metallic elements crystallize in one of three simple crystal structures, in which each atom is symmetrically surrounded by eight or twelve nearest neighbors. Metals can form compounds with each other in which the bonding is mainly metallic. These intermetallic compounds exhibit a wide variety of structural types. Except for hydrogen, all the elements in columns 1-12 of the periodic table are metals. Columns 13-16 include some metals, some non-metals, and several elements with intermediate properties. Al, Ga, Sn, Bi, and Po are predominantly metallic, while elements above or to the right of them have progressively fewer metallic properties. Some compounds between metals and non-metals, for example the carbides and nitrides of the metals in columns 4-6, are predominantly metallic.

The energies involved in all three primary forms of chemical bonds are on the order of one to ten electron volts per atom. While the loss or gain of a single electron is easiest for atoms with only one electron or one vacancy in an unfilled shell, multiply-charged ions are more strongly attracted to each other in a crystal lattice. For example, sodium fluoride (Na^+F^-) melts at 1269 K and boils at 2075 K, while magnesium oxide ($Mg^{2+}O^{2-}$), which has the same crystal structure and the same total number of electrons, melts at 3105 K. Atoms whose unfilled shells are more nearly half-filled can form greater numbers of covalent bonds or form stronger metallic bonds than those in the outer columns of the periodic table. The trends in melting points (see Figure 6, page 15) of the elements within each row of the periodic table illustrate this effect.

Atoms that have already attained a filled outer shell by electron transfer or sharing can still participate in secondary forms of attraction. One of the stronger forms of secondary attraction is called hydrogen bonding. A hydrogen atom that is already covalently bonded to a highly electronegative atom (F, O, or N) in one molecule can form a hydrogen bond to an F, O, or N atom in another molecule. Hydrogen bonding is largely responsible for the high boiling point of H_2O in comparison to many other small covalent molecules. Molecules with polar covalent bonds have non-uniform distributions of electrical charge. The more negative portion of a polar molecule will be attracted to positive ions or to the positive portions of other polar molecules.

Even nonpolar molecules and single neutral atoms are subject to secondary attractive forces. The electric field of a nearby ion or polar molecule can perturb the electron cloud of a nonpolar molecule or atom, inducing a small temporary charge redistribution that results in a weak attraction. Finally, the electron clouds of nonpolar molecules or atoms can perturb each other enough to generate weak attractive forces. The secondary attractions that affect nonpolar molecules and atoms become stronger with increasing size of their electron clouds. This trend is illustrated by the boiling points of the elements in column 18, which vaporize as single atoms, and those in column 17, which form covalently bonded diatomic molecules.

Secondary attractions would not be important on a hot planet, but on the earth's surface they cause water and all but the smallest organic molecules to be liquids or solids rather than gases. They cause some (but not all) ionic solids to dissolve in water. They hold together the two halves of the double helix of DNA, the molecule that carries the genetic code, and they cause protein molecules to fold themselves into characteristic three-dimensional structures. Determining the secondary structures of proteins experimentally and predicting them theoretically are among the most important current challenges for structural chemistry, and for molecular biology.

[130] I. Langmuir, J. Amer. Chem. Soc. 41 868, 1543 (1919).
[131] Linus Pauling, *The Nature of the Chemical Bond. Third Edition*, (Cornell University Press, Ithaca, NY, 1960).

Overview of the Chart of the Nuclides

Earlier Charts and Tabulations

As new nuclides were produced using techniques described in the previous section, charts and tables summarizing properties of the known nuclides were developed. One of the earliest charts was published in 1935 by Giorgio Fea[39] and a portion of that chart is shown in Figure 5 on page 13. Charts with other formats were used by researchers to emphasize different properties or relationships between the nuclides.[132, 133] Emilio Segre published a more detailed chart, similar to current charts, in 1946.[134] In Segre's chart, masses, abundances, capture cross sections, half-lives and decay modes were given for the known nuclides. As with Fea's chart, the neutron number was used as the vertical axis and proton number as the horizontal axis.

In 1957, a trilinear chart of the nuclides was published by W.H. Sullivan.[135] A portion of the chart is shown in Figure 9. Here, each nuclide is represented by a hexagon. Within the hexagons, half-lives, spins, decay modes, thermal neutron cross sections and other information about each nuclide were given. With this arrangement, **isobars** (nuclides having the same mass number) lie in a vertical line; **isotopes** (nuclides with the same atomic number) lie along a line inclined at 30° above the horizontal; **isotones** (nuclides with the same number of neutrons) lie along a line inclined at 30° below the horizontal; and **isodiapheres** (nuclides having the same excess number of neutrons, A-2Z) lie on a horizontal line. Additional information about the trilinear coordinate system is given in Reference 133.

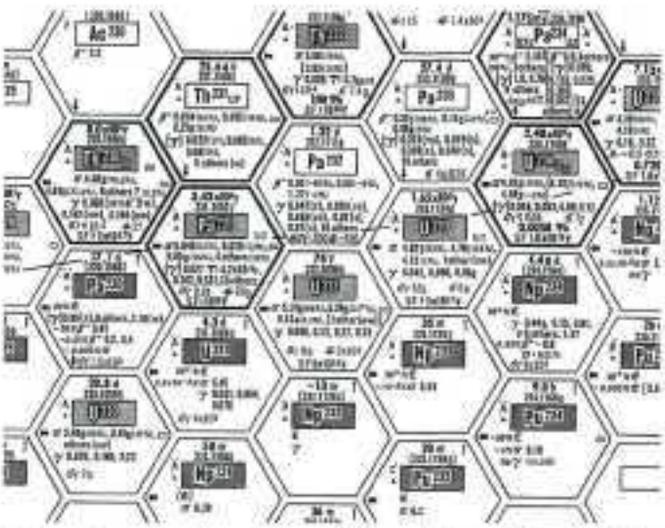


Figure 9. A Portion of the Trilinear Chart of the Nuclides Published by W.H. Sullivan (see Ref. 135).

One of the first comprehensive tabulations of properties of known nuclides was published by Livingood and Seaborg in 1940.[136] This tabulation was the first edition of the Table of Isotopes, now in its eighth edition.[137]

The first edition of the present Chart of the Nuclides was compiled by G. Friedlander and M. Perlman and published by the General Electric Company in 1946. This Chart differed from those of Fea and Segre, in that the proton number was used as the vertical axis and the neutron number as the horizontal. All editions of the Chart of the Nuclides since then have used this layout. The current chart is the Sixteenth Edition of the Chart of the Nuclides.

At present, there are about 3100 known nuclides and 580 known isomers (see Isomeric States, page 21). Except for short-lived isomers present in decay of longer-lived parents or as part of the naturally-occurring radioactive series, only isomers with half-lives greater than about one second are included in the above count. Active nuclear research, which is conducted in many laboratories throughout the world, causes additions and changes in the list of nuclides. A given nuclide is included on the Chart only if the work leading to its discovery has been published in a refereed, scientific journal. Proceedings of technical conferences and the most recent edition of the *Table of Isotopes*[137] are also consulted for the evaluation. Data referenced as "private communication" are not included in the Chart. Since the last edition was published (1996), approximately 300 new nuclides and 100 new isomers have been added to the Chart. In addition, there has been at least one change in more than 95% of the squares on the Chart.

General Layout of the Chart

Information about the known nuclides, both stable and unstable, is summarized on the Chart of the Nuclides. Each nuclide occupies a square in a grid where the atomic number, Z, is plotted against the number of neutrons, N. The atomic number is given at the left of each horizontal row. Each row represents one element; the gray squares indicate the known stable isotopes (and in some cases, long-lived naturally-occurring isotopes) of that element. Each vertical column shows nuclides (isotones) having the same neutron number, N.

The heavily bordered space at the left side of each horizontal row gives the name and properties of the element. Each of the other occupied spaces in the row carries the chemical symbol and the mass number of the nuclide indicated. The mass number, designated by A, is the sum of the number of neutrons, N, and protons, Z, in the nucleus.

Nuclides on diagonals running from the upper left to the lower right on the Chart have the same mass numbers (isobars). For example, one diagonal could connect ^{40}Ca, which has 20 protons and 20 neutrons, with ^{40}Ar, which has 18 protons and 22 neutrons.

Heavy lines on the Chart occur for N or Z equal to 2, 8, 20, 28, 50, and 82, and for N = 126. These are the so-called "magic numbers", i.e. the numbers of neutrons (protons) present when a neutron (proton) shell is closed. (see The Nuclear Shell Model, page 27).

Colors and shading are used on the chart squares to indicate relative magnitudes of half-lives and neutron absorption properties. Four different colors—blue, green, yellow, and orange—are used for these identifications. The background color of the upper half of a square represents half-life and the lower half is representative of the greater of the thermal neutron cross section or resonance integral. An exception is made for fully gray-shaded squares, which simply extends the gray "stable nuclide" shading of the upper square throughout, when the cross section is small or unknown.

The gray shading is also used for unstable nuclides having lifetimes sufficiently long (>5x10^8 years or 5E8 years, see Footnote 138) to have survived from the time the elements were formed. These naturally-occurring radioactive nuclides are highlighted by a special black band or rectangular area near the

[132] See Clark Goodman, Ed., *Introduction to Pile Theory*, (Addison-Wesley Press Inc., Cambridge, Mass., 1952).
[133] D.E. Hull and W.H. Sullivan, J. Chem. Education **24** 436 (1947).
[134] Emilio Segre, *A Mind Always in Motion, The Autobiography of Emilio Segre*, (University of California Press, Berkeley, 1993), page 190. Segre's original chart was developed at Los Alamos in 1945. An unclassified revision of the first chart was available in 1946. A third revision, carried out in 1948, is included in Reference 132.
[135] W.H. Sullivan, *Trilinear Chart of Nuclides, 2nd Edition*, (U.S. Government Printing Office, Washington, D.C., 1957).
[136] J.J. Livingood and G.T. Seaborg, Rev. Mod. Phys. **12** 30 (1940).
[137] R.B. Firestone, S.Y. Frank Chu and C.M. Baglin, *Table of Isotopes, Eighth Edition, 1998 Update*, (John Wiley and Sons, New York, 1998).

[138] Throughout the Chart and this booklet, the designation E is used to mean power of ten, e.g. 5E8 = 5x10^8.

top. Other black-banded nuclides are formed via natural processes such as cosmogenically in the atmosphere (^{3}H, ^{7}Be, and ^{14}C) or are daughter products of naturally-occurring radioactive parents. The nuclides belonging to the naturally-occurring radioactive decay series are indicated by smaller black rectangles containing the symbol for the historic name, e.g. RaA, MsTh$_1$, and Io, given when these series were first investigated. White or colored squares represent artificially produced radioactive nuclides.

Certain squares, e.g. ^{60}Co, ^{115}In, and ^{116}In, are divided. Such divisions occur when a nuclide has one or more isomeric (or metastable) states. **Isomers** are nuclides having the same A and Z, but different radioactive properties, i.e. different energy states of the same nucleus.

Fission products from the slow neutron fission of ^{233}U, ^{235}U, or ^{239}Pu are indicated by small black triangles or squares in the lower right corner of the squares. The known isobaric fission chains are indicated on the Chart by lines drawn diagonally off the corner of the first or last nuclide for each mass number A. For a given value of A, the total cumulative yield for ^{235}U thermal neutron fission is given at the lower end of the diagonal line; the total cumulative yields for ^{233}U and ^{239}Pu thermal neutron fission are given at the upper ends, with the ^{239}Pu yields in parentheses.

Question marks are used to indicate uncertainty in nuclide identification or mode of decay.

Information about the products of various nuclear reactions can be easily obtained from the Chart of the Nuclides. Figure 10 illustrates the combinations of incident (or bombarding) and emitted particles for several types of reactions and how each combination changes the original nucleus. A special type of shorthand (see Nuclear Reactions, page 12) is used on this diagram to identify the reactions. An example is (p,n) which denotes a reaction in which the nucleus absorbs a proton and emits a neutron. For this reaction, the product nucleus would have the same value of A as the target, but the value of Z would have increased by one. On the Chart, the product nucleus is then located one square diagonally upward and to the left of the target nucleus. For the (n,α) reaction, the target nucleus would have lost two protons and one neutron, resulting in a product nucleus two squares down and one square to the left of the target.

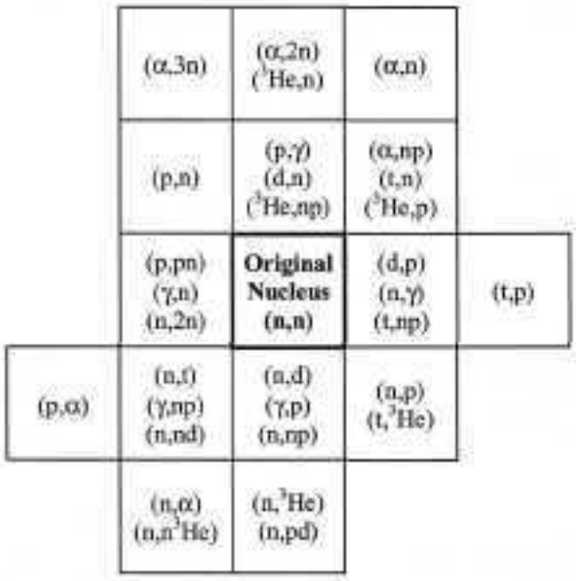


Figure 10. Changes Produced by Various Nuclear Reactions. *See page 22 for symbol definitions.*

There is not enough space to indicate the one standard deviation uncertainty for the numerical values on the Chart. Therefore, the following convention has been adopted: the values given have been rounded off so that the uncertainty in the last decimal place is less than or equal to five units. For example, a value of 6.78±0.08 barns would be quoted as 6.8. Similarly, a half-life value of 2500±300 years would be quoted as 2.5E3. The symbol, "~", is used to show data having errors greater than the limit of five units in the last figure. This is commonly used when only one measurement with an uncertainty greater than five has been reported in the literature. Thus, a cross section of 2.7 ± 0.8 barns may be quoted as ~2.7. If trailing zeros are followed by a decimal point, the zeros are all significant. Otherwise, the last non-zero digit is significant and has an uncertainty less than or equal to 5.

Properties of the Nuclides Displayed on the Chart

The type of information given in each square of the Chart depends on whether the square corresponds to a chemical element, a stable nuclide or a radioactive nuclide. Several example squares are shown on the Chart, with an indication of the type of information given in each of the squares. Table 2 on page 19 lists the properties for which values (when known) are given for each type of square on the Chart. Each property is described in detail in Sections A-K, below.

A. Chemical Element Names and Symbols

The element names and symbols used on the Chart are those recommended by the International Union of Pure and Applied Chemistry (IUPAC). These include the revised recommendations for the transfermium elements with atomic numbers 101 through 109 and new recommendations for 110.[119] Elements above Z = 110 have not been named officially by IUPAC at the time of publication and are listed solely by atomic number.

B. Spins and Parities

The upper right corner of the square for the ground state of the nuclide and the upper left corner of the isomeric state contain the spin and parity of the corresponding energy level. These are depicted as J^{π}, where J represents the spin (either integral or half-integral) and π represents the parity (either "+" or "-"). Each neutron and proton has an intrinsic spin of ½ (in units of ħ), similar to that of the electron. The nucleon spin combines with the orbital angular momentum to produce a resultant angular momentum called the nuclear spin. Since the orbital angular momentum is always zero or an integral multiple of ħ, the nuclear spin (in units of ħ) is always integer or half-odd-integer, depending upon whether the nucleus has an even or odd number of nucleons. The concept of parity was introduced by the mathematical formalism of quantum theory and has no classical analogue. For ^{27}Al, the spin and parity are shown on the Chart as 5/+, where the 2 in the denominator of 5/2 has been removed to improve the readability. The ground states of all even-even nuclides are known to have spin and parity 0^{+}, so 0^{+} has been omitted from the Chart squares for these nuclides.

The arguments for the assignment of spin and parity to nuclear states can be divided into two classes: strong arguments such as measuring values directly and weak arguments such as inferring values indirectly. On the Chart, the absence of parentheses indicates spins and/or parities based on strong arguments; the presence of parentheses indicates spins and/or parities based on weak arguments. An effort has been made to not include those values inferred solely from systematics.

C. Isotopic and Atomic Masses

Isotopic masses in atomic mass units (u) are given on the Chart for stable isotopes, for certain long-lived naturally-occurring radioactive isotopes (see Table 3, page 24), and for those nuclides where particle (alpha, proton or neutron) decay becomes a prominent mode (>10%) of decay. If the nuclide

[119] G.P. Moss, Pure Appl. Chem. 69 2471 (1997). See also http://www.iupac.org/reports/provisional/abstract03/corish_300603.html.

Table 2. Types of Information Given for Elements and Nuclides on the Chart of the Nuclides[141]		
Type of Square	**Property**	**See**
1. Chemical Element	a. Chemical symbol. b. Atomic mass on the $^{12}C = 12.000\,000$ mass scale. c. Chemical element name. d. Total thermal neutron absorption cross section and absorption resonance integral in barns per atom.	a. Section A b. Section C c. Section A d. Section G
2. All Nuclides	a. Chemical symbol and mass number. b. Spins and parities of the ground and metastable states, except for even-Z even-N nuclides. c. Fission products produced in the thermal neutron fission of ^{235}U, ^{233}U, and ^{239}Pu are indicated.	a. Section A b. Section B c. Section K
3. Stable Nuclides	a. Isotopic abundance, on an atom percent basis. b. Isotopic mass of the neutral atom, on the $^{12}C = 12.000\,000$ mass scale. c. Thermal neutron cross sections and resonance integrals for significant absorption reactions. The values of these quantities determine the background color of the lower half of the chart square. d. Also see "All Nuclides."	a. Section D b. Section C c. Section G
4. Radioactive Nuclides	a. Half-life. The magnitude of the half-life determines the background color of the upper half of the chart square. b. Major modes of decay and decay energies (gamma-ray energies in keV, all other particle energies and decay energies in MeV). c. An order of magnitude indication of absolute abundance of radiation emitted when less than one percent. d. Beta disintegration energies in units of MeV and/or the isotopic mass of the neutral atom on the $^{12}C = 12.000\,000$ mass scale. e. Isomeric (Metastable) States. f. Long-lived radioactive nuclides and members of the naturally-occurring radioactive decay series are indicated by a black band near the top of the square. g. If measured, the thermal neutron cross sections and resonance integrals for significant absorption reactions. The values of these quantities determine the background color of the lower half of the chart square. h. Also see "All Nuclides."	a. Section E b. Section H c. Section H d. Sections I and C e. Section F f. Section J g. Section G

decays primarily by β^+, ε, or β^- decay, the disintegration energy (see Section I) is given instead. If particle decay and either β^+, ε, or β^- decay are both prominent (> 10%) modes of decay, both the isotopic mass and the total disintegration energy are given. Isotopic masses were taken from either the 1995 study of G. Audi and A.H. Wapstra[140] or more recent references (see Primary Technical References (Chart of the Nuclides), page 37).

The chemical atomic masses (relative atomic mass of an element) listed on the Chart are those promulgated by the Commission on Atomic Weights—Inorganic Chemistry Division of the International Union of Pure and Applied Chemistry (IUPAC) in its 2001 report.[142] The uncertainties in the atomic masses are given in the tables on the inside front cover and on page 48 of this booklet. Occasionally, a sample can display a significant difference from the accepted atomic mass because of artificial isotope fractionation, artificial nuclear reaction, or a rare geological occurrence in a small quantity. (See Isotopic Abundances, below.)

For most elements, the uncertainty in the isotopic composition has an effect on the atomic mass which is at least two orders of magnitude greater than the uncertainties in the isotopic masses. Differences in the number of significant figures given on the Chart for the elemental and isotopic masses of mono-isotopic elements are the result of the different estimates of uncertainties given by Audi and Wapstra and the Commission on Atomic Weights, respectively.

D. *Isotopic Abundances*

The isotopic abundance values on the Chart are given in atom percent and were taken from Rosman and Taylor.[143] A few isotopic abundance values were quoted in which the uncertainty in the last figure is greater than 5. This was done so that the isotopic abundance values total exactly 100% for each element.

Isotopic abundances are specified on the Chart for 288 nuclides. These include 266 stable and 22 radioactive isotopes. Of the radioactive isotopes, 20 are long-lived with half-lives greater than 5E8 years (about a tenth of the age of the earth[144]) and are listed in Table 3 on page 24. The remaining two, ^{231}Pa ($t_{1/2}$ = 3.28E4 years) and ^{234}U ($t_{1/2}$ = 2.46E5 years), are continually being formed through the decay of ^{235}U and ^{238}U, respectively (see Figure 12, page 25).

Isotopic Abundance Variability. Unlike the mass of an individual nuclide, the atomic mass of an element having more than one stable isotope is not an intrinsic property of that element. It depends on the relative abundance of the different isotopes. The different isotopes are produced by different processes, at different places in the universe, and at different times. The earth is as homogeneous as it is with respect to isotopic abundance because it is so small compared to interstellar distances.

Some of the largest natural abundance variations on the earth are displayed by elements that have one or more isotopes that can be formed by decay of a natural radioisotope with a half-life in the range of $\sim 10^7$-10^{11} years (see Table 3, page 24). Lead, which has isotopes formed by decay of both thorium and uranium, has already been mentioned. Helium and argon are two other prominent examples. The original atmosphere of the earth is believed to have had isotopic compositions similar to the sun, but most of this early atmosphere was apparently lost to space and replaced later by outgassing from the condensed earth. In the solar system as a whole, ^{36}Ar is the dominant argon isotope. The earth's atmosphere has the same $^{38}Ar/^{36}Ar$ ratio as the sun, but a much greater $^{40}Ar/^{36}Ar$ ratio, because the argon remaining on the earth is dominated by ^{40}Ar from decay of ^{40}K. The helium isotopic composition on the Chart refers to helium in the earth's

[140] G. Audi and A.H. Wapstra, Nucl. Phys. **A595** 409 (1995).
[141] X-ray energies are not included on the Chart. However, intensity-averaged X-ray energies are given on the Periodic Table on the inside back cover of this booklet.
[142] T.B. Coplen, Pure Appl. Chem. **73** 667 (2001).
[143] K.J.R. Rosman and P.D.P. Taylor, Pure Appl. Chem. **70** 217 (1998).
[144] C.J. Allegre, G. Manhes, and C. Gopel, Geochim. Cosm. A., **59** 1445 (1995)

atmosphere. The atmospheric $^4He/^3He$ ratio is higher than the sun's, due to loss of much of earth's primordial ^3He, and replacement by ^4He produced by alpha decay of heavy elements. Helium in old uranium-rich crustal rock has an even higher $^4He/^3He$ ratio than the atmosphere, while rocks recently derived from mantle sources have lower $^4He/^3He$ ratios than the atmosphere. Among the mantle-derived rocks, those with the lower $^4He/^3He$ ratios are believed to come from sources deeper in the mantle.[145,146] Strontium, neodymium, and hafnium also display variations in isotopic abundance, caused by decay of rubidium, samarium, and lutetium, respectively. Xenon isotope abundance is affected by spontaneous fission of actinide elements (including ^{244}Pu) and decay of primordial ^{129}I. Anomalous Mg isotope abundances, resulting from decay of ^{26}Al, have been found in some of the oldest meteorites (those that solidified earliest during the formation of the solar system).[147,148,149]

Natural fractionation processes can also cause variations in isotopic abundances, especially for the lighter elements. Deuterium-to-hydrogen ratios for terrestrial samples vary by up to 70%.[150] Fresh water, ice, and snow are depleted in deuterium relative to seawater. Oxygen (once used as a standard of atomic mass) is fractionated, but less so, by the same mechanisms that fractionate hydrogen. Compared to natural water, atmospheric O_2 has a higher $^{18}O/^{16}O$ ratio.[151]

Human technology can also cause variation in isotopic abundances, intentionally or unintentionally. Purification by distillation can cause greater variations in isotopic abundance for elements such as calcium than are found naturally.

On the shelf, chemical reagents with a 3.75% ^6Li composition, instead of the 7.59% given on the Chart have been noted without any warning of the depletion in ^6Li content.[143] Natural variations in boron from 19.1% to 20.3% ^{10}B have been measured. The atomic masses and the thermal neutron absorption cross sections for those materials will then differ significantly from the values listed on the Chart. When atomic masses or neutron absorption cross section values are critical to an application, it is highly recommended that samples be checked for actual isotopic abundance before use.

The precision with which variations in isotopic abundance can be measured has improved greatly over the years, and the necessary sample size has decreased. A large and growing community of scientists uses these variations to study subjects ranging from earth's former climates[152, 153] to ancient trade routes.[154] The community has developed standard analysis procedures and standard reference materials for calibration. The books by E. Durrance[155] and by G. Faure[159] provide good introductions to the entire field.

E. Half-lives

The half-life is the period of time in which half of the nuclei initially present in a given radioactive sample disintegrate. On the Chart, half-lives for radioactive nuclides are given below the nuclide symbol and mass number. The units used for half-lives are:

μs	microseconds (1.0E-6 s)
ms	milliseconds (1.0E-3 s)
s	seconds
m	minutes
h	hours
d	days
a	years

A number of so-called particle unstable nuclides with extremely short half-lives are included on the Chart. All of these nuclides are located in the low mass region below A = 10, e.g. ^5He, ^5Li, ^6Be, ^8Be, and ^9B. The rationale for including these nuclides is that they can be formed in nuclear reactions involving nearby stable nuclides and resonances corresponding to the formation of these compound nuclei have been observed in the reactions. The half-lives of these nuclides cannot be measured directly, as they are too short to be determined with currently available techniques. Rather, the half-lives are inferred from the measured level widths of the resonances observed in the reactions using the Heisenberg uncertainty principle,

$$\Delta E \Delta t = \hbar$$

Background Color of the Upper Half of the Chart Square. For nuclides with half-lives of one day or longer, the background color in the upper half of the Chart square indicates the relative magnitude of the half-life of the nuclide. The following ranges of half-lives are used:

1 day to 10 days—orange
>10 days to 100 days—yellow
>100 days to 10 years—green
>10 years to 5E8 years—blue

The color gray represents stable or naturally-occurring radioactive nuclides—the 288 nuclides for which isotopic abundances have been assigned by IUPAC (see Reference 143). For the 266 stable nuclides and for the 20 long-lived ($t_{1/2}$ > 5E8 years) radioactive nuclides, listed in Table 3 (page 24), the background color in the upper half of the square is gray. For ^{231}Pa and ^{234}U, with half-lives less than 5E8 years, a gray band is used at the top of the square, while the color in the remainder of the upper half corresponds to the half-life of the nuclide.

Half-Life Variability. Half-lives for beta decay, electron capture, or conversion electron emission depend on the local electron density as well as on the properties of the nucleus. Over the full range of astrophysical conditions, the variations in half-lives can be enormous, but under ordinary terrestrial conditions the variations are so small that extraordinary care is needed to measure them experimentally. Low electron density slows electron capture and accelerates β^- decay or conversion electron emission (see Section H). Fully ionized ^7Be nuclei, as cosmic rays, travel for many years without undergoing electron capture because there are no nearby electrons. Fully ionized ^{163}Dy nuclei have been observed in the laboratory to emit β^- particles with a half-life of 48 days resulting[158] in the production of ^{163}Ho. Under normal terrestrial conditions, the opposite process occurs—electron capture by ^{163}Ho to produce ^{163}Dy. At the opposite extreme, in dense stars the electron density is high enough to drive electron capture reactions that would not occur on earth. This process of forcing electrons into nuclei, converting protons into neutrons, culminates in the formation of a neutron star.

Under terrestrial conditions, variations in half-life depending on chemical environment are generally less than one percent, and can easily be swamped by experimental artifacts. Nevertheless, careful work has shown that real variations do exist. In contrast to isotopic abundance variations, the techniques for measurements of half-life variations have improved little over the years. New half-life variation reports appear only sporadically and are seldom verified by independent measurements in another

[145] C.J. Allegre, T. Staudacher, P. Sarda, and M. Kurz, Nature **303** 762 (1983).
[146] P.J. Tackley, Science **288** 2002 (2000).
[147] C.M.O'D. Alexander, A.P. Boss, and R.W. Carlson, Science **293** 64 (2001).
[148] A. Gaily, E.D. Young, R.D. Ash and K. O'Nions, Science **290** 1751 (2000).
[149] K.D. McKeegan, M. Chaussidon and F. Robert, Science **289** 1334 (2000).
[150] G. Faure, *Principles of Isotope Geology, 2nd Edition*, (John Wiley and Sons, 1986), p. 430.
[151] *ibid.*, p. 436.
[152] C.F.T. Andrus, D.E. Crowe, D.H. Sandweiss, E.J. Reitz, and C.S. Romanek, Science **295** 1508 (2002).
[153] E.J. Hendy *et al.*, Science **295** 1511 (2002).
[154] G. Giuliani *et al.*, Science **287** 631 (2000).
[155] E.M. Durrance, *Radioactivity in Geology*, Ellis (Harwood Ltd., Chichester UK, 1986).
[158] M. Jung *et al.*, Phys. Rev. Lett. **69** 2164 (1992); O. Klepper, Nucl. Phys. **A626** 199 (1997).

laboratory. For more information, the reader is referred to the review article by H. Daniel.[157]

The disintegration constants of individual nuclear states are not affected by temperature, but different nuclear states have different decay properties, and nuclides can change states through interaction with photons.[158] Only photons of high energy are able to induce nuclear transformations. On earth, there are too few ambient high-energy photons to have any effect on nuclear decay. At more extreme temperatures, though, more high-energy photons become available. In stars, the decay of ^{26}Al depends greatly on temperature because of a low-lying excited state that has a β^+ decay half-life more than twelve orders of magnitude shorter than that of the ground state.

F. Isomeric (Metastable) States

Certain squares on the Chart are divided, e.g. the square for ^{26}Al. Such divisions occur when a nuclide has one or more isomeric states—when nuclides have the same mass numbers and atomic numbers but possess different radioactive properties in different long-lived energy states. An isomeric state is included on the Chart if the half-life is one second or longer. This arbitrary cutoff point of one second was chosen for convenience in presenting the data. Shorter half-lives are included when they are fed in the decay of a parent nuclide with half-life longer than a second or are part of a naturally-occurring radioactive decay series. Recent research near the proton drip line (see page 31) has allowed the identification of low-lying isomeric states with half-lives (~1 μs) and particle energies significantly different but close to those of the ground state of the nuclide. These isomeric states are included on the Chart to better understand the data used to calculate the ground state half-lives and particle decay energies in this region. When referring to isomeric states, the lower energy state is generally referred to as the ground state and the higher energy state as the isomeric state. Frequently, the ground state is a stable nuclide. If one metastable state exists, it is shown on the left of the square and the ground state is to the right. If two exist, the higher energy state is shown on the upper left of the square, the lower below it and the ground state to the right of both. When a metastable and ground state (or two metastable states) cannot be unequivocally assigned to a given energy level, the designation ↔ is placed between the two states involved.

When the spins of both the ground state and an isomeric state are given for a particular nuclide, it is interesting to observe that these spins usually differ by two or more units of $\hbar$. A large angular momentum (spin) change is required for a gamma-ray transition between the states. Combining this spin change with the small energy differences (a few hundred keV) leads to a relatively long lifetime (of the metastable state).

G. Neutron Cross Sections and Resonance Integrals

The neutron cross section of a nuclide is a measure of the probability of the interaction of a neutron with the nuclide and is a function of the neutron energy. The cross section can be most easily visualized as a cross-sectional target area presented to the neutron by the nucleus. Cross sections are usually measured in units of barns per atom. A barn is 10^{-24} square centimeters.

The Greek letter σ, with various subscripts is used to denote cross sections for different types of reactions. Several types of reactions of neutrons with nuclei are possible. The neutron can scatter from the target nucleus, leaving the nucleus either unchanged (elastic scattering) or in an excited state (inelastic scattering). The neutron can also be "absorbed" in reactions with the target nucleus. The most probable absorption reaction (the reaction with the largest cross section) is, generally, the neutron capture reaction, in which the absorption of the neutron by the nucleus is accompanied by high-energy gamma-ray emission; the capture cross section is indicated by the symbol σ_γ. Occasionally, a proton or an alpha particle may be emitted, or a nucleus may fission upon neutron absorption and the cross sections for these reactions are indicated by σ_p, σ_α, and σ_f, respectively. Examples of these cross sections are found on the squares of ^{7}Be, ^{10}B, and ^{227}Th. The neutron absorption, or disappearance, cross section (symbol σ_a) is the sum of the cross sections for all absorption reactions:

$$\sigma_a = \sigma_\gamma + \sigma_p + \sigma_\alpha + \sigma_f + \ldots$$

For a parallel beam of neutrons, or neutron flux, of φ neutrons per cm^2 per second, the reaction rate per atom (reactions per second per atom) is given by the product of the cross section (cm^2) and the flux, $\sigma\varphi$. When the neutron flux is not a monodirectional beam, the interaction rate is still given by the product $\sigma\varphi$, where φ is the product of the neutron density (cm^{-3}) and the neutron speed (cm/sec).

Because much of the data given on the Chart has been obtained for the purpose of, or as the result of, the development of fission reactors, the cross sections and neutron energy spectra are presented in that context. Neutrons produced in the fission process have an average energy of approximately 2 MeV (2E6 eV). In a thermal reactor these neutrons are slowed down by collisions with the moderator atoms. Therefore, a spectrum of neutron energies exists in a reactor. An arbitrary division frequently used for this neutron energy spectrum is comprised of three regions:

1. The thermal region, in which neutrons are in thermal equilibrium with the moderator.
2. The intermediate or slowing down region in which the neutron spectrum varies inversely with neutron energy due to energy loss by elastic scattering.
3. The fast region in which the neutron spectrum is governed by the yield from the fission process.

When neutrons reach thermal equilibrium with the moderator, their energies are determined by the thermal energy distribution of the moderator atoms and the neutron energy spectrum becomes a Maxwellian distribution at the temperature (T in kelvin) of the moderator material. Setting the kinetic energy of the neutron motion equal to the thermal energy of the moderator, one can obtain the most probable speed (magnitude of the velocity) of the neutrons in thermal equilibrium. For a moderator at room temperature (20°C), this neutron speed is 2200 meters per second and the corresponding energy is 0.0253 eV. Neutron cross sections specified for this energy are called thermal neutron cross sections.

In the thermal region, the absorption cross section of many materials varies as $E^{-1/2}$, where E is the neutron energy. These materials are called 1/v absorbers. At low energies, the ^{6}Li(n,α), ^{10}B(n,α), ^{197}Au(n,γ), ^{235}U(n,f), and ^{238}U(n,γ) are a few of the reactions which have this energy dependence.

In the intermediate energy range, there are often particular energies for which the cross sections are exceptionally high, especially for nuclides with intermediate and high mass numbers. Graphs of the cross sections of these materials versus energy in this region show narrow peaks, called resonances, at incident neutron energies corresponding to the discrete energy levels in the compound nucleus of the reaction.

A quantity called the resonance integral (RI) is used to represent the probability of neutron reactions in the energy range above thermal energies. In order to evaluate this integral over a typical neutron slowing down spectrum, where resonance absorption peaks occur, a 1/E function is used to represent the neutron spectrum. The range of integration is from the cutoff energy of the thermal range, E_c, typically about 0.5 eV, to the highest energy where appreciable contributions to the reaction rate occur.

[157] H. Daniel, At. Energy Rev. **17(2)** 287 (1979).
[158] R.A. Ward and W.A. Fowler, Astrophys. J. **238** 266 (1980).

$$RI = \int_{E_c}^{E_{max}} \sigma(E) \frac{dE}{E}$$

In the present work, the absorbing nuclide is assumed to be present in such dilute quantities that there is no perturbation of the neutron slowing down spectrum. Thus, the resonance integral is called an infinitely dilute resonance integral.

Unlike the thermal cross section, the resonance integral is a reaction rate integral over a broad energy range. To derive an average cross section for the energy range, the resonance integral is divided by the integral of dE/E over the same energy range.

$$\overline{\sigma_{res}} = \frac{RI}{\int_{E_c}^{E_{max}} \frac{dE}{E}}$$

For historic reasons, and because the magnitudes of thermal cross section, σ_{th}, and resonance integral, RI, are similar, these two quantities are given for each nuclide to represent the reaction probabilities for the thermal and above thermal ranges even though the resonance integral is not itself a cross section. In the remainder of this discussion and in common practice, the resonance integral is treated as a cross section.

The measurement of the reaction rate in a reactor neutron spectrum produces contributions from intermediate energy neutrons with a 1/E distribution and from thermal neutrons with a Maxwellian distribution of energies. To distinguish between these two contributions, a cadmium cover is often used to filter out the thermal neutrons. Cadmium filters do not separate the thermal from the intermediate neutron spectrum sharply, but an "effective cadmium cutoff energy", E_{cd}, can be defined for a 1/v absorber. Suitable cadmium filters have an E_{cd} of approximately 0.5 eV. This energy is sufficiently high to exclude most of the low-energy deviations of the flux from a 1/E shape. The lower limit on the integral, above, defining the resonance integral is then taken as 0.5 eV.

On the Chart the thermal (2200 m/s value) cross section for a particular reaction is listed as σ with a subscript appropriate for that reaction. This 2200 m/s value is listed if it has been directly measured (e.g. by time-of-flight methods) or if it is inferred from integral measurements relative to a standard. A cross section for neutrons with energies below the cadmium cutoff is listed as $\bar{\sigma}$ with subscripts. All other cross sections including reactor spectrum values are shown as $\dot{\sigma}$ with subscripts.

Updates to the thermal cross sections and resonance integrals reported in the Fifteenth Edition of the Chart of the Nuclides were made based on information obtained from CINDA 2000.[159] Data were obtained from either the EXFOR[160] files at the National Nuclear Data Center (NNDC) at Brookhaven National Laboratory or technical publications. Comparisons with other compilations[161] of thermal cross sections and resonance integral data were carried out.

For elemental materials the absorption cross section is given on the leftmost square of each row. For a stable or long-lived nuclide, cross sections for specific, significant, reactions, e.g. capture, fission, etc., are given. In all cases, the thermal cross section symbol is given, followed by the cross section value and the resonance integral value, separated by a comma. The designation **mb** or **μb** following the cross section value indicates that the units of the cross section or resonance integral is millibarns per atom or microbarns per atom, respectively. When no **mb** or **μb** appears on the Chart square, the units of the cross section are barns per atom. A given nuclide might undergo two or more types of reactions and its square would then contain two or more of these cross section values.

Background Color of the Lower Half of the Chart Square. The background color in the lower half of the Chart square is determined by the greater of the neutron cross section or resonance integral. Four colors—blue, green, yellow and orange—are used to indicate ranges of values from 10 barns to greater than 1000 barns. The following ranges of neutron absorption properties are used:

10 to 100 barns—blue
>100 to 500 barns—green
>500 to 1000 barns—yellow
>1000 barns—orange

When neutron capture can lead to a metastable state as well as to the ground state, more than one value will appear beside the capture cross section symbol for that nuclide. The cross section value for metastable-state formation is listed on the left and that for direct-ground-state formation on the right followed by similar values for the resonance integral. For two metastable states, the higher of the two states is on the left. For example, ^{115}In has an indicated cross-section and resonance integral of (87 + 75 + 41), (2.6E3 + 6.7E2). This means that the thermal cross section for formation of the 2.16 second state of ^{116}In is 87 barns, the thermal cross section for the direct formation of the 54.2 minute ^{116}In is 75 barns and the resonance integral for direct plus indirect formation (by decay of the 2.16 second state) is 2.6E3 barns. The thermal cross section and resonance integral for formation of the 14.1 second ground state of ^{116}In are 41 barns and 6.7E2 barns, respectively. When the ground state cross section is not known but one of the metastable cross sections is, a question mark is used as a placeholder to designate this case (see ^{124}Sn).

II. Major Modes of Decay and Decay Energies

The modes of decay indicated on the Chart include:

α	alpha particle (^{4}He nucleus)
β^-	beta minus (negative electron)
β^+	beta plus (positron)
γ	gamma-ray
n	neutron
p	proton
d	deuteron (^{2}H nucleus)
t	triton (^{3}H nucleus)
ε	electron capture
IT	isomeric transition
e^-	conversion electron
SF	spontaneous fission
$\beta^-\beta^-$	double beta decay
⊗	cluster decay
D	delayed radiation

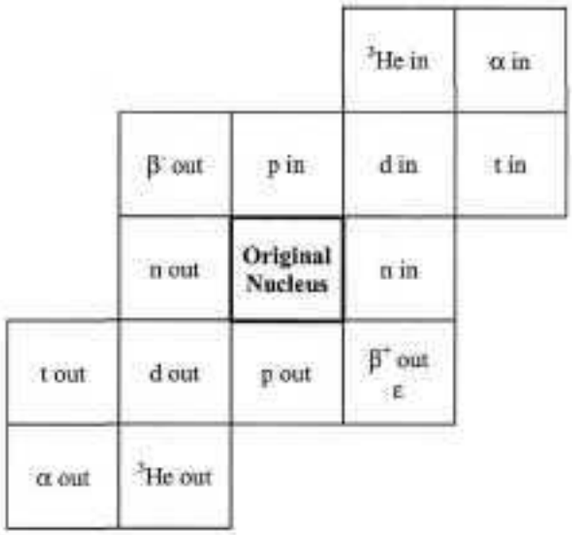
As these processes occur, whether in natural radioactivity or under artificially induced conditions, the nuclides change in accordance with the scheme shown in Figure 11, page 23. To understand the use of this scheme more fully, consider the decay of ^{238}U. As noted on the Chart, ^{238}U decays by the emission of an alpha particle. The daughter nucleus is in the second space diagonally down to the left (see Figure 11). This square represents the isotope ^{234}Th. ^{234}Th is also radioactive and decays by β^- emission. In this decay the mass number, A, does not change; however, the loss of one negative charge means that the atomic number Z increases by one. In effect, one neutron in the nucleus has changed to one proton. The move one space up and one space to the left (see Figure 11) leads to ^{234}Pa which has

[159] International Atomic Energy Agency, *CINDA 2000, The Index to Literature and Computer Files on Microscopic Neutron Data*, (IAEA, Austria, Nov. 2000). Also on the Internet at http://www.nndc.bnl.gov/nndc/cinda/.

[160] Experimental Nuclear Reaction Data File (EXFOR[CSISRS]), National Nuclear Data Center, Brookhaven National Laboratory. On the Internet at http://www.nndc.bnl.gov/nndc/exfor.

[161] R. Kinsey, "The NuDat Program for Nuclear Data on the Web", NNDC, Vers. 2.5, Aug. 1996, on the Internet at http://www.nndc.bnl.gov/nndc/nudat/xriform.html.
N.E. Holden, "Neutron Scattering and Absorption Properties", in *CRC Handbook of Chemistry and Physics, 81st Edition, op. cit.*

isomeric states. Each of these states undergoes β^- emission, so another move diagonally upward to the left leads to ^{234}U. The effects of other modes of decay—β^+, ε, n, p, etc.—on the atomic number and neutron number can be seen in Figure 11.



n = neutron	α = alpha particle
p = proton	β^- = beta minus (negative electron)
d = deuteron	β^+ = beta plus (positron)
t = triton	ε = electron capture

Figure 11. Relative Locations of the Products of Various Nuclear Processes on the Chart of the Nuclides.

A metastable (or isomeric) state of a nuclide will frequently decay to the ground state by **Isomeric Transition (IT)** gamma-ray emission, followed by one or more gamma-rays in cascade. It is possible to have a more complicated decay scheme with more than one IT gamma-ray, each with its cascading gamma-rays as well as a simple decay scheme with only one IT gamma-ray to the ground state of the daughter nucleus. In order to indicate which gamma-rays are leaving the metastable level, the IT gamma-rays are listed separately from the cascading gamma-rays. For example, ^{84}Rb has a metastable state 463.6 keV above the ground state. This metastable state can decay directly to the ground state by emitting a single 463.6 keV gamma ray. It can also decay to a state of intermediate energy by emitting a 215.6 keV gamma ray. The intermediate state has a half-life of less than a nanosecond and decays by emitting a 248.0 keV gamma. In the ^{84}Rb square on the Chart, the IT gamma energies, 463.6 keV and 215.6 keV, are listed after "IT", while the cascade gamma energy, 248.0 keV, is listed on a separate line.

Internal conversion, a process resulting from the interaction between the nucleus and its extranuclear electrons, competes with gamma-ray emission. Here the excitation energy of the nucleus is transferred to an orbital electron, usually K-shell, and the electron is emitted from the atom instead of a gamma-ray. Neither Z nor N of the nucleus change in this process. The conversion electrons have kinetic energies equal to the difference between the energy of the nuclear transition involved and the binding energies of the electrons in the atom. Since internal conversion is an alternative to gamma-ray emission, the intensity of a given gamma-ray may be reduced considerably, with the difference in intensity from that expected accounted for by the conversion electron intensity. The total internal conversion coefficient, α_T, is the ratio of the total number of internal conversion electrons to the number of gamma-rays emitted for a specific transition. It is the policy for the Chart to indicate conversion electrons (symbol **e**$^-$) only if α_T is equal to or greater than 10. This arbitrary cutoff for inclusion of conversion electrons on the Chart is a compromise between significant effect on the gamma-ray intensity and the desire to increase the legibility of the Chart. Occasionally, the symbol **e**$^-$ is included in parentheses to indicate that it is specific to a given, listed gamma-ray.

Internal conversion coefficients vary over a wide range. Internal conversion is favored by high Z, low transition energy, and large angular momentum (spin) change. For example, the 156.0 keV transition of ^{117m}Sn results from the decay of the 11/2- state to an intermediate 3/2+ state. The total internal conversion coefficient, α_T, for this transition is 46, so it is indicated with an **e**$^-$ on the Chart. The cascading 158.6 keV gamma results from decay of the 3/2+ state to the 1/2+ ground state. For the 158.6 keV transition, α_T is only 0.16, so it is not followed by **e**$^-$ on the Chart. As an extreme example, the 28.6 keV transition of ^{113m}Ag from the (7/2+) state to the (1/2-) ground state has an internal conversion coefficient of 12,000. (When the **e**$^-$ symbol is used, ω or other symbols indicating low photon abundance are sometimes omitted.)

Spontaneous fission (SF) is a decay mode[162] of some heavy nuclei where the nucleus breaks up into two intermediate-mass fragments and several neutrons. It has been observed only for nuclei with A > 230. See, for example, the squares corresponding to ^{234}U, ^{238}Pu or ^{242}Cm on the Chart. This mode of decay is discussed further on page 30.

Double beta decay ($\beta^-\beta^-$) is a very rare radioactive event, studied[163] first in 1935 and observed[164] for ^{92}Mo and ^{100}Mo in 1987. In this process, two beta particles and two antineutrinos are generated, resulting in the original nucleus gaining two protons and losing two neutrons. Several nuclides, including ^{82}Se, ^{116}Cd, and ^{130}Te, have been observed to undergo this type of decay, with half-lives greater than 10^{19} years. Decay by emission of two betas and no neutrinos may also be possible according to some theories that require neutrinos to have non-zero mass. Experiments are in progress to try to detect zero-neutrino double beta decay, but as yet no positive evidence has been found.[165]

Cluster decay has been observed in several heavy nuclides. In this type of decay, the emission of ^{12}C, ^{14}C, ^{20}O, ^{24}Ne, ^{28}Mg or ^{32}Si clusters has been observed. Specific nuclei that undergo cluster decay, the emitted cluster and further information are listed in Table 5, page 31. Isotopes that decay by cluster radioactivity are denoted on the Chart by the symbol

located in the lower right hand corner of the square. More information on this mode of decay is given on page 31.

Very often a nuclide decays initially to an excited state of the daughter nuclide. The most common mode of decay for daughter excited levels is gamma emission. When the daughter level has too short a half-life to have its own place on the Chart (or has a half-life much shorter than that of the parent nuclide), the gamma energy is listed with the parent nuclide. However, if the daughter level has a half-life greater than approximately one microsecond, the symbol **D**, meaning **delayed radiation,** is placed after the gamma energy. An example is provided by 30.07-year ^{137}Cs. This long-lived parent decays by β^- emission to a short-lived daughter, 2.552-minute ^{137}Ba. The 661.7 keV gamma-ray, which is emitted by the ^{137}Ba, is included on the ^{137}Cs square with a "**D**" to indicate delayed radiation proceeding through an isomeric level of the daughter nucleus. As another example, the three gamma-rays listed below the β^- symbol for the 2.9 minute state of ^{138}Cs are all emitted from excited states of ^{138}Ba. The 191.9 keV gamma-ray is marked with a **D**, because it is emitted by a state with a 0.8 microsecond half-life, while the other two gammas are

[162] Spontaneous fission of uranium was first observed by G.N. Flerov and K.A. Petrzhak, J. Phys. U.S.S.R. 3 275 (1940). The emission of neutrons with spontaneous fission was first reported by G. Scharff-Goldhaber and G.S. Klaiber, Phys. Rev. **70** 229 (1946).
[163] M. Goeppert-Meyer, Phys. Rev. **48** 512 (1935).
[164] S.R. Elliott, A.A. Hahn and M.K. Moe, Phys. Rev. **36** 2129 (1987).
[165] V.I. Tretyak and Y.G. Zdesenko, At. Data Nucl. Data Tables **80** 84 (2002).

emitted by states with half-lives much shorter than a microsecond.

Delayed particle emission can also occur since excited states of daughter nuclides can decay by modes other than gamma emission. A mode of decay shown in parentheses indicates that the decay results from emission of the indicated particle from an excited level of the daughter nucleus, following positive or negative beta-decay to that level. For example, ^{17}N, with a half-life of 4.174 seconds, decays by β^- emission into several exceedingly short-lived states of ^{17}O, which in turn emit neutrons. This is indicated on the Chart by the symbol **(n)**, followed by the energies of the most abundant neutrons, if they have been measured. Thus, nitrogen-17 emits "delayed-neutrons" with a half-life of 4.174 seconds. Another example is ^{23}Si, with a half-live of 220 milliseconds, which decays by β^+ emission to short-lived states in the daughter nucleus ^{23}Al. These excited states, in turn, emit protons. The "delayed-protons" are indicated in the Chart by the symbol **(p)**. Cases are also known where "delayed-alphas" are emitted, indicated by the symbol **(α)**. Delayed emission of two protons is possible after the first beta-decay, and this process is indicated by the symbol **(2p)**. The emission of two particles following beta decay also occurs with delayed neutron-emission, **(2n)**, and delayed alpha-emission, **(2α)**. The energy requirement for the occurrence of delayed particle emission is that the energies of the levels in the daughter nucleus reached by the beta emission exceed the separation energy of the proton(s), neutron(s), or alpha(s) from this nucleus. Beta decay of some of the heaviest nuclides leads to excited daughter levels that decay by spontaneous fission. This is indicated on the Chart by the symbol **(SF)**.

The location of modes of decay, specific particle energies, or gamma-ray energies within a square conveys information about the **abundances or intensities**. Different modes of beta decay (ε, β^+, or β^-) appear on separate lines if the intensity of one of the modes of decay is less than 10% absolute intensity. Conversely, they appear on the same line if the intensity of both is greater than 10% absolute intensity, with the most abundant listed first. In addition, the designations "ω" (weak), "υω" (very weak), and "υυω" (very very weak) are used to indicate particle and gamma-ray energies or decay modes with less than one percent, less than 10^{-3} percent, and less than 10^{-4} percent absolute intensity, respectively. Of course, if one of a series of particle or gamma-ray energies has one of the above designations, then all succeeding energies have this designation also (since these energies are arranged in order of decreasing intensity). The designations "ω", "υω", and "υυω" are often placed in parentheses for clarity of expression. All cluster decay observed is "υυω"— see Table 5, page 31.

Numeric decay information given on the Chart is obtained from a comprehensive literature search of published results as well as from other tabulations of this type of information.[137, 166]

I. Disintegration Energies

If the nuclide decays primarily (>10%) by β^+, ε, or β^- decay, the disintegration energy, denoted by "E", is given. Where both β^+ and β^- decay occur, the respective decay energies are denoted by E+ and E-. The disintegration energy refers to the total energy difference between the ground state of a parent radionuclide and the ground state of its daughter. The energy given is the maximum possible energy for β^-, but for β^+ decay, 1.022 MeV ($2m_ec^2$) must be subtracted to get the maximum possible β^+ particle energy. In beta decay, the total energy is shared between the beta particle and the neutrino or antineutrino.

J. Naturally-Occurring Radioactive Isotopes

Sixty-five naturally-occurring radioactive isotopes are currently known. On the Chart these naturally-occurring radioactive nuclides are highlighted either by a black band at the top of the square or, for nuclides forming the naturally-occurring radioactive decay chains, a smaller black rectangular area near the top of the square in which the symbol corresponding to the historic name of the nuclide is given. These isotopes generally belong to one of three classes:

1. *Cosmogenically produced nuclides present in the atmosphere.* This class of naturally-occurring radioactive materials includes ^3H, ^7Be, and ^{14}C. These isotopes are all formed in the atmosphere at concentrations greater than 0.01 picocuries per kilogram of air by reactions initiated by cosmic rays. Both ^3H and ^{14}C decay by β^- emission and have half-lives of 12.32 and 5715 years, respectively. Both have been used to determine the ages of objects that contain them. ^7Be decays by electron capture and has a half-life of 53.3 days. Filters through which air has passed are radioactive due to both ^7Be and radon chain products, which deposit on dust and airborne particles.

2. *Unstable nuclides having lifetimes sufficiently long to have survived from the time the elements were formed.* Materials belonging to this class, along with their half-lives (in years), and abundances, are listed in Table 3. These isotopes decay by β^+, ε, β^-, or α decay and, with the exception of ^{232}Th, ^{235}U, and ^{238}U, form stable products. The three nuclides, ^{232}Th, ^{235}U, and ^{238}U, are the parents of naturally-occurring radioactive decay series and are discussed below. As noted in the Table, in addition to the very long half-lives, many materials in this group have very low abundances, resulting in low activities and making them difficult to detect. As experimental techniques improve, other isotopes may be added to this class of naturally-occurring radioactive materials.

A few nuclides, for example, ^{82}Se, ^{116}Cd, and ^{130}Te, have been found to decay by $\beta^-\beta^-$ emission with half-lives greater than 10^{19} years. Because of the unusual type of decay and extremely low activities, these nuclides were not included in Table 3.

3. *Naturally-occurring radioactive decay chains.* The last three nuclides listed in Table 3 all decay to shorter-lived radioactive daughters forming long chains or radioactive series. Because the parents are found in nature, their daughters achieve a significant

Table 3. Long-Lived Naturally-Occurring Radioactive Isotopes.

isotope	decay	half-life (years)	isotopic abund. (%)	decay product
^{40}K	β^-, β^+	1.27 E09	0.0117	^{40}Ca, ^{40}Ar
^{50}V	ε, β^-	1.4 E17	0.250	^{50}Ti, ^{50}Cr
^{87}Rb	β^-	4.88 E10	27.83	^{87}Sr
^{113}Cd	β^-	7.7 E15	12.22	^{113}In
^{115}In	β^-	4.4 E14	95.71	^{115}Sn
^{123}Te	ε	6 E14	0.89	^{123}Sb
^{138}La	ε, β^-	1.05 E11	0.090	^{138}Ba, ^{138}Ce
^{144}Nd	α	2.38 E15	23.80	^{140}Ce
^{147}Sm	α	1.06 E11	14.99	^{143}Nd
^{148}Sm	α	7. E15	11.24	^{144}Nd
^{152}Gd	α	1.1 E14	0.20	^{148}Sm
^{176}Lu	β^-	3.75 E10	2.59	^{176}Hf
^{174}Hf	α	2.0 E15	0.16	^{170}Yb
^{180m}Ta	ε, β^+	>1.2 E15	0.012	^{180}Hf
^{187}Re	β^-	4.12 E10	62.60	^{187}Os
^{186}Os	α	2. E15	1.59	^{182}W
^{190}Pt	α	6.5 E11	0.014	^{186}Os
^{232}Th	α	1.40 E10	100.	^{208}Pb
^{235}U	α	7.04 E08	0.720	^{207}Pb
^{238}U	α	4.47 E09	99.2745	^{206}Pb

[166] J.K. Tuli, Ed. Nucl. Data Sheets, through June 2002.

equilibrium concentration where the parent activity is found. The three series are named the thorium, uranium, and actinium series, after elements at, or near, the head of the series. Each of these series ends at a stable isotope of lead—^{206}Pb, ^{208}Pb, or ^{207}Pb. These series are shown in Figure 12. Excluding the stable end products and the parent nuclides, these three series contain a total of 42 radioactive nuclides.

The thorium series has ^{232}Th ($t_{1/2}$ = 1.40E10 years) as the parent and after a series of alpha and β^- emissions, reaches ^{208}Pb, a stable end product. In many of the steps in the series, gamma-rays and conversion electrons are also emitted. The uranium series has ^{238}U ($t_{1/2}$ = 4.47E9 years) as the parent and ^{206}Pb as the end product. The actinium series has ^{235}U ($t_{1/2}$ = 7.04E8 years) as the parent and ^{207}Pb as the end product. Collateral series, i.e. series running parallel to one of the naturally-occurring series have also been investigated. For example, the nuclide ^{227}Pa decays by successive alpha emissions to ^{207}Pb. This series is comparable to the actinium series.

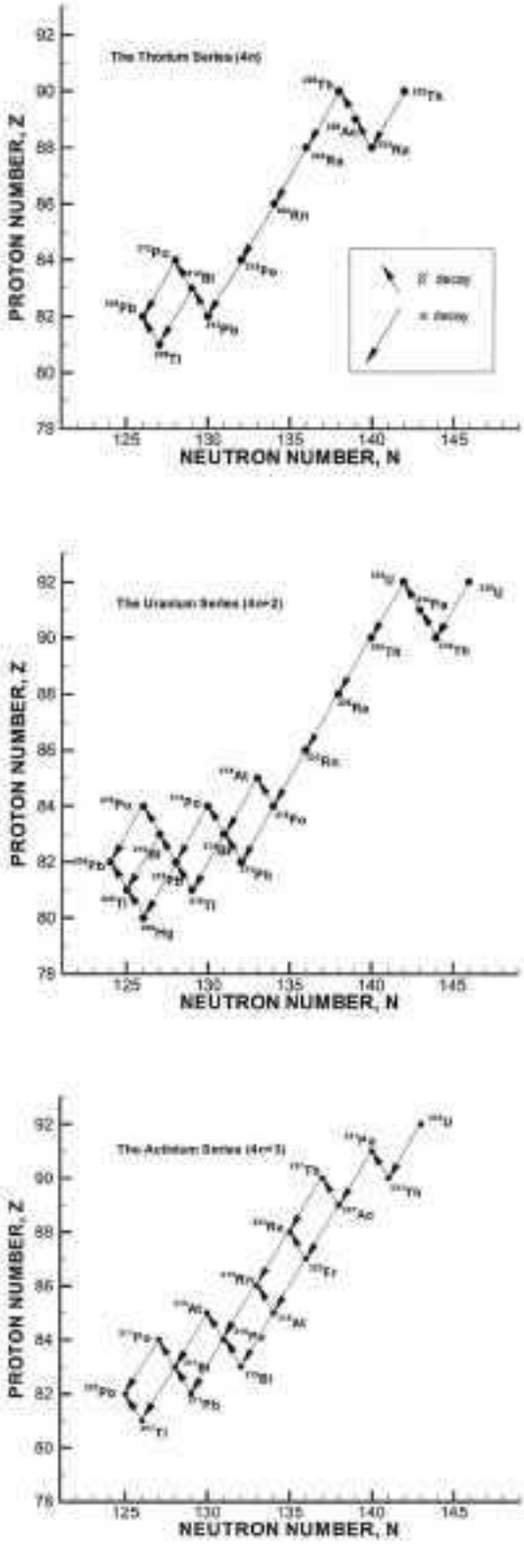


Figure 12. The Three Naturally-Occurring Radioactive Decay Chains.

The historical symbols, e.g. RaA for ^{218}Po, are given on the Chart for members of these chains. These names were assigned in the early 1900s (see "Radioactivity", page 6) while the series were being investigated. The symbols usually indicate the chain to which a given isotope belongs: RaA, RaB,..., RaG all belong to the uranium series; AcA, AcB,..., AcD belong to the actinium series; ThA, ThB,..., ThD belong to the thorium series. The relative position of an isotope in the chain can usually be determined from the historic name, i.e. "A" decays to "B", which decays to "C", etc. In addition, similarly named isotopes of the three series are isotopes of the same element: RaA, AcA, and ThA are all isotopes of Po; RaB, AcB, and ThB are all isotopes of Pb.

As noted above, the members of these series decay by α or β^- emission. Since decay by alpha emission decreases the atomic mass, A, of a nuclide by 4 and beta emission has no effect on the value of A, the mass numbers of members of a given series differ by multiples of 4. Accordingly, the A value of any member of the thorium, uranium, or actinium series can be characterized by the expression 4n, 4n + 2, or 4n + 3, respectively, where the n's are integers ranging from ~50 to 59.

A radioactive series of nuclides whose mass numbers are represented by 4n + 1 does not exist in nature. This series, the neptunium series, beginning with ^{241}Pu and ending with ^{209}Bi, has been studied. The longest-lived member of this 4n + 1 series is ^{237}Np, with a half-life of 2.14E6 years. Since the age of the earth is about 5E8 years, most of the ^{237}Np present when the earth was younger has already decayed and the series is not found in nature.

It should be noted that the end products of the naturally-occurring decay series—^{206}Pb, ^{207}Pb and ^{208}Pb—are all "magic" nuclides (see The Nuclear Shell Model, page 27) with Z = 82. ^{208}Pb is a "doubly magic" nuclide with N = 126. The end product of the neptunium series, ^{209}Bi, is also magic with N = 126.

K. Fission Product Designations and Fission Yields

Binary fission of ^{233}U, ^{235}U or ^{239}Pu produces many different fission products. Fission products range from nickel (Z = 28) to dysprosium (Z = 66), and from mass number A = 72 to A = 163 in the thermal neutron fission of ^{233}U. Fission products from the slow neutron fission of ^{235}U, ^{233}U, and ^{239}Pu are indicated on the Chart by small black triangles or squares (see below) in the lower right corner of the squares.

Since the neutron-to-proton ratio is about 1.6 for ^{235}U and about 1.3 for the stable nuclides in the fission product region, the primary products of fission are on the neutron-rich side of stability. Each such product decays by successive β^- decays to a stable isobar. No neutron-deficient nuclides have been found among the products of thermal neutron fission; however, a few so-called shielded nuclides occur among the fission products. A shielded nuclide has a stable isobar one unit lower in Z so that it is not formed as a daughter product in a beta-decay chain. Examples are ^{80}Rb and ^{136}Cs. The fission yield of such a nuclide is presumably due entirely to its direct formation as a primary product (independent yield). However, its presence in fission products could also be due to neutron capture by the stable fission product to its left on the Chart. Shielded yields greater than 2.5E-6% are indicated by a black square (■) in the lower right-hand corner of the Chart squares for those nuclides.

The independent yield of a particular fission product is the direct or instantaneous yield of that nuclide without any contribution from decay of preceding members of the isobaric

chain. The measured total cumulative yield of a particular fission product represents the sum of its independent yield and of the independent yields of all its precursors.

The known isobaric fission chains are indicated on the Chart by lines drawn diagonally off the corner of the first or last nuclide square for each mass number A. For a given A, near the end of the diagonal line extending downward from the isobar with lowest Z is the total cumulative yield for ^{235}U thermal neutron fission; at the end of the diagonal line extending upward from the isobar with the largest Z are total cumulative yields for ^{233}U thermal neutron fission (above) and ^{239}Pu thermal neutron fission (below in parentheses). The total cumulative yield of a nuclide is the percentage of the total fissions that leads directly or indirectly to that nuclide. Since two fission products are emitted for every binary fission, the total cumulative yields of all the isobaric chains add up to 200%. The fission yields given on the Chart of the Nuclides are total cumulative chain yields (in percent) of the last isobaric chain member of ^{235}U, ^{233}U, and ^{239}Pu thermal neutron fission. The yields are from the work of England and Rider.[187]

The fission yields from the thermal neutron fission of ^{235}U, ^{239}Pu, and ^{233}U are plotted against mass number, A, in Figure 13. Fission into two equal fragments is by no means the most probable mode in thermal neutron fission. As can be seen in the figure, asymmetric modes are much more favored, with the maximum yields occurring at A = 95 and A = 138 for ^{235}U thermal neutron fission. The asymmetry becomes less pronounced with increasing bombarding energy of the neutron.

Ternary fission, or breakup into three products, occurs less frequently than binary fission. This process is a source of light charged particles such as ^{3}H, ^{3}He, ^{4}He, ^{7}Be, etc. These long-lived or stable, light nuclides can build up in the nuclear fuel of an operating reactor.

An enormous amount of radiochemical and mass spectrometry work was required to arrive at our present state of knowledge about fission products. It was necessary to develop chemical separation procedures, analyze radioactive decay and growth patterns, determine beta- and gamma-ray energies, establish mass assignments of many previously unknown nuclides, and measure the fission yields. This work continues at many universities and national laboratories throughout the United States and the world.

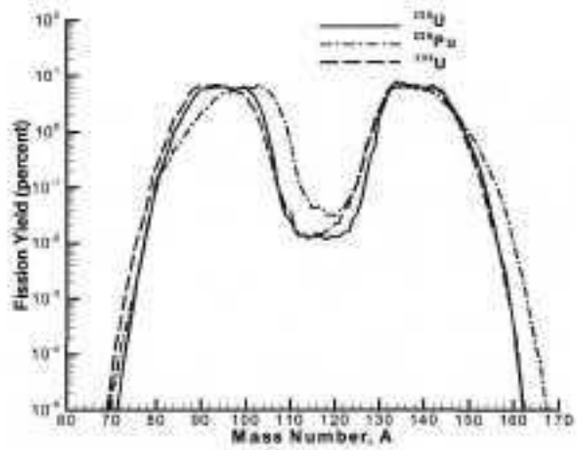


Figure 13. Comparison of the Thermal Fission Yields for ^{235}U, ^{239}Pu and ^{233}U.

[187] T.R. England and B.F. Rider, "Evaluation and Compilation of Fission Product Yields", LA-UR-94-3106, Los Alamos National Laboratory, Oct. 1994.

Trends in Stability on the Chart

Stable Nuclides

As shown on the Chart, 266 stable nuclides (gray colored squares) exist. Figure 14 is a plot of the neutron number, N, versus the proton number, Z, for these nuclides. Also shown on the graph is the Z = N line. The deviation of the (N,Z) values of the stable nuclei from the Z = N line indicates the number of neutrons in these nuclei is always greater than or equal to the number of protons for Z > 2. The light stable nuclides cluster around the Z = N line, but heavier stable nuclides have many "excess" neutrons. For these nuclides, as Z increases, the long-range Coulomb repulsion between protons is balanced by the presence of additional neutrons providing additional short-range attractive nuclear forces.

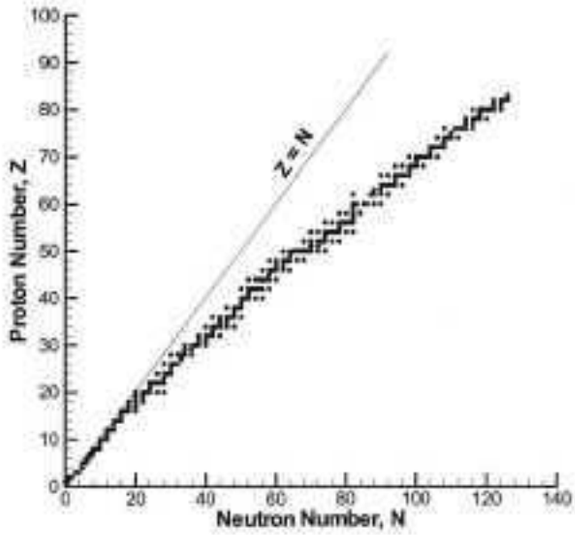


Figure 14. Values of (N,Z) for the 266 Stable Nuclides.

Classifying the stable nuclides by the evenness or oddness of Z and N gives four possible categories. The first category contains an even number of protons and an even number of neutrons (so called even-even nuclei). The other categories are even-odd, odd-even, and odd-odd. Table 4 shows the number of stable nuclides that fall in each category.

As seen in Table 4, there are more stable nuclides with even Z and even N than in any other category. More stable nuclides with even Z exist than with odd Z; more stable nuclides exist with even N than with odd N. For the odd A nuclides there are approximately as many nuclides with an even number of protons (even Z) as with an even number of neutrons (even N). This is evidence that the nuclear force between two nucleons is independent of whether the nucleons are protons or neutrons. Odd-odd stable nuclides are scarce, being found only among the lightest nuclides.

Table 4. Distribution of Stable Nuclides.

A	Z	N	Number of Stable Nuclides
Even	Even	Even	159
Odd	Even	Odd	53
Odd	Odd	Even	50
Even	Odd	Odd	4
			266

The Nuclear Shell Model

The periodicity of many chemical properties of elements led to the development of the shell theory of the electronic structure of atoms. Similarly, studies[168] of properties of nuclides–the number of stable isotopes of each element, the relative natural abundances of isotopes, variations of the magnitude of the thermal neutron capture cross section for different nuclides–led to the development of a nuclear shell model for the neutrons and protons within a nucleus.[169] As pointed out earlier, isotopes (and isotones) with Z (or N) equal to the "magic numbers" 2, 8, 20, 28, 50, and 82, and N = 126 are indicated on the Chart by heavy lines. These values correspond to filled proton or neutron shells and represent the most stable configurations. There are more stable isotopes (or isotones) for nuclides having these values of N or Z than for other values. For example tin, Z = 50, has ten stable isotopes. Several nuclides, e.g. ^{4}He, ^{16}O, ^{40}Ca, ^{48}Ca, and ^{208}Pb, have both N and Z values that are magic numbers and are referred to as "doubly magic."

The scarcity of odd-odd stable nuclides is due to a "pairing energy" between particles in the same shell. Nucleons residing in the same shell increase the binding energy of these particles, making them more stable than particles in different shells. An odd-odd nuclide contains at least one unpaired proton and one unpaired neutron, which usually reside in different shells and hence contribute weakly to the binding. However, for the lightest nuclei, ^{2}H, ^{6}Li, ^{10}B and ^{14}N, the unpaired neutron and proton are in the same shell.

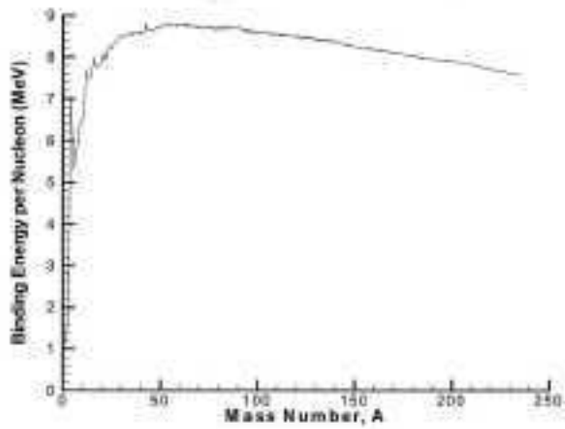


Figure 15. Binding Energy per Nucleon as a Function of Mass Number for the Stable Nuclides.

Binding Energy

The difference between the actual nuclear mass of a nuclide and the mass of all the individual nucleons is called the total binding energy. It represents the work necessary to dissociate the nucleus into separate nucleons or, conversely, the energy which would be released if the separated nucleons were reassembled into a nucleus. For convenience, the masses of atoms, rather than the masses of the nuclei are used in all calculations. The mass, M(A,Z), of an atom having Z protons and (A-Z) neutrons in the nucleus is given by

$$M(A,Z) = ZM_H + (A-Z)M_n - B(A,Z)$$

[168] O. Haxel, J.H.D. Jensen and H.E. Suess, Phys. Rev. 75 1766 (1949); M.G. Mayer, Phys. Rev. 78 1969 (1949).

[169] see M. G.-Mayer and J.H.D. Jensen, *Elementary Theory of Nuclear Shell Structure*, (John Wiley and Sons, New York, 1955) and references within.

where M_H is the mass of the hydrogen atom (including the electron), M_n is the neutron mass, and B(A,Z) is the binding energy.[170] Since B(A,Z) is positive, a stable atom has a smaller mass than the combined masses of the constituent particles. The binding energy per nucleon, B(A,Z)/A, for the stable nuclides, is plotted as a function of A in Figure 15, page 27. As seen in the figure, the binding energy per nucleon rises sharply, reaching a maximum value of about 8.8 MeV near A = 50 and then slowly decreases to a value of ~7.6 MeV near A = 230. Below A = 20, the three small spikes in the curve correspond to ^{4}He, ^{12}C and ^{16}O. The binding energies per nucleon of these three nuclides are considerably larger than neighboring nuclides—for ^{4}He, it is approximately 7.1 MeV.

The fact that the curve in Figure 15 peaks at A~50 indicates that the binding energy could be increased by either breaking heavy nuclides apart (fission) or by fusing light nuclides together (fusion). Both processes would result in the release of energy.

In a nuclear reaction, the difference in the rest masses of the initial particles and final products determines whether the reaction is endothermic or exothermic. For the reaction

$$X(x,y)Y$$

the Q-value of the reaction is defined as

$$Q = (M_X + M_x) - (M_Y + M_y).$$

If Q is positive for the reaction, the reaction is exothermic and energy is released during the reaction. If Q is negative, the reaction is endothermic and energy is required for the reaction to take place.

For ^{235}U, the fission reaction initiated by neutrons can be written as:

$$^{235}U + n \rightarrow {}^{236}U^* \rightarrow X + Y + \bar{\upsilon}\, n$$

where X and Y are two intermediate mass nuclides and $\bar{\upsilon}$ is the average number of neutrons released. The distribution of masses for thermal neutron fission of ^{235}U is shown in Figure 13 (page 26). For nuclides in the vicinity of uranium, the binding energy per nucleon is about 7.6 MeV, while it is about 8.5 MeV in the vicinity of the fission products (A ~ 70 to A ~ 160). This is about 0.9 MeV/nucleon greater in the fission products. The total energy released in the fission reaction is, then, in excess of 200 MeV. The fission fragments are neutron rich (see below) and decay predominately by β^- emission. The average number of neutrons, $\bar{\upsilon}$, emitted in the thermal neutron fission of ^{235}U is approximately 2.43 and it is the emission of these neutrons that makes a chain reaction possible. Most are "prompt" neutrons, emitted at the time of fission, but a small fraction are emitted as "delayed" neutrons from specific fission products. The latter are very important in the control of nuclear reactors.

For small A, the binding energy per nucleon rises rapidly with mass number as shown in Figure 15. In fusion reactions, light nuclides are combined to form heavier nuclides. The products are more tightly bound and energy is released in the reaction. Some important examples of fusion reactions are:

$$^2H + {}^2H \rightarrow {}^3H + p \quad (Q = 4.03 \text{ MeV})$$
$$^2H + {}^2H \rightarrow {}^3He + n \quad (Q = 3.27 \text{ MeV})$$
$$^2H + {}^3H \rightarrow {}^4He + n \quad (Q = 17.6 \text{ MeV})$$
$$^2H + {}^3He \rightarrow {}^4He + p \quad (Q = 18.4 \text{ MeV})$$

While the energy released in these fusion reactions is much less than that released in a fission reaction, the energy yield per unit mass of material is greater in fusion than in fission.

[170] The equation for M(A,Z) neglects the binding energy of the electron in the hydrogen atom and the binding energy of all electrons in the atom specified by A and Z. These energies are small compared to other terms in the equation.

Unstable (Radioactive) Nuclides

Figure 16 shows a graph of the values of (N,Z) for all nuclides given on the Chart. The filled squares indicate stable nuclides, while the open squares correspond to the unstable or radioactive nuclides. For a given A, the binding energy B(A,Z) is largest for the stable nuclides and decreases for both larger and smaller values of Z. In the figure, those nuclides with corresponding squares lying above the stable nuclides have too many protons for stability (proton-rich), while those with squares lying below the stable nuclides have too many neutrons for stability (neutron-rich).

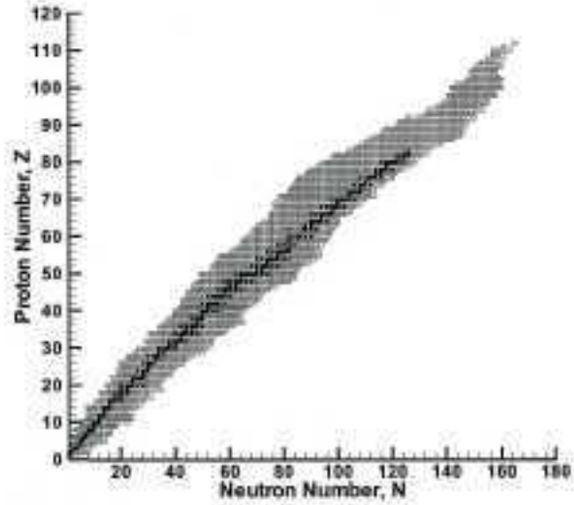


Figure 16. Plot of Neutron Number versus Proton Number for all Known Nuclides. *Filled squares represent stable nuclides and open squares correspond to unstable nuclides.*

If a third dimension, mass, were added to the graph in Figure 16, the resulting surface would resemble a valley with the stable nuclides along the bottom of the valley. Rising on one side of the valley would be the proton-rich nuclides and on the other would be the neutron-rich nuclides. These unstable nuclides on the sides of the valley undergo radioactive decay bringing the final product closer to a stable configuration, moving towards the bottom of the valley. The modes of radioactive decay to achieve this differ, depending on the mass region of the Chart.

Beta Decay and Electron Capture

Below Z~83, radioactive nuclides move towards stability by increasing or decreasing the nuclear charge through beta decay (β^+, β^-) or electron capture (ε). In these processes either a positive or negative electron is emitted from the nucleus or an orbital electron is captured by the nucleus. Nuclides having an excess of neutrons (those below the stable nuclides shown in Figure 16) tend to undergo β^- decay. For β^- decay, a neutron within the nucleus decays as:

$$n \rightarrow p + \beta^- + \bar{\nu}$$

where $\bar{\nu}$ is an antineutrino. In this process, the atomic number of the initial nuclide increases by one, while A is unchanged:

$$^A_Z X \rightarrow {}^A_{Z+1} Y + \beta^- + \bar{\nu}$$

In terms of atomic masses, the condition for β^- decay to be possible is:

$$M(A,Z) \geq M(A,Z+1)$$

Fission products are neutron-rich nuclides and tend to undergo successive β^- decays forming the decay chains described earlier.

Nuclides having an excess of protons (those above the stable nuclides in Figure 16, page 28) tend to undergo β^+ decay or electron capture. For β^+ decay, a proton in the nucleus decays as

$$p \rightarrow n + \beta^+ + \nu$$

where ν is a neutrino. In this process, the atomic number of the initial nuclide is decreased by one:

$$^{A}_{Z}X \rightarrow {}^{A}_{Z-1}Y + \beta^+ + \nu$$

In terms of atomic masses, the condition for β^+ decay is

$$M(A,Z) \geq M(A,Z-1) + 2m_e$$

where m_e is the mass of an electron.[171]

In electron capture, ε, an atomic electron is captured by a proton in the nucleus:

$$p + e^- \rightarrow n + \nu$$

resulting in the atomic number of the original nuclide decreasing by one. In terms of atomic masses, for electron capture to be possible

$$M(A,Z) \geq M(A,Z-1)$$

Note that β^+ decay is possible only if the initial and final masses differ by $2m_e$ (1.022 MeV), while electron capture, ε, is possible if the initial mass is just greater than the final mass. If β^+ decay is possible, electron capture is also. Nuclei that cannot undergo β^+ decay can, however, undergo decay by electron capture.

In the mass region where beta decay is the dominant mode, there are no more than three stable isobars for a given A. The number of stable isobars is determined by the oddness or evenness of A.

Figure 17 presents a plot of the masses of isobars with A = 61 as a function of Z. For this value of A, only one stable nuclide exists—^{61}Ni with Z = 28. Nuclides with larger Z values (proton-rich) decay by β^+ emission while nuclides with smaller values of Z (neutron-rich) decay by β^- emission. Both types of decay modes move towards stable ^{61}Ni. Generally, for a given odd value of A (Z-odd, N-even or Z-even, N-odd), a plot of the masses of the isobars versus the atomic number gives a parabola-like curve with the most stable isobar at the bottom of the curve. Typically, for a specific odd A only one stable isobar exists.

For even values of A, two or three stable isobars are typically found in nature. Even A nuclides have either Z and N odd, or Z and N even. Masses of the isobars with A = 40 are plotted in Figure 18. The upper curve connects the masses of the odd-odd nuclides, while the lower curve connects the masses of the even-even nuclides. Two stable nuclides, ^{40}Ca and ^{40}Ar, exist for this value of A; both are even-even. The odd-odd nuclide, ^{40}K, lies between them. It decays either by β^+ emission to ^{40}Ar or by β^- emission to ^{40}Ca.

The Chart lists 33 nuclides that undergo both β^+ and β^- decay, e.g. ^{36}Cl, ^{40}K, ^{70}Ga, ^{74}As. These are all odd-odd nuclides and are located between beta-stable, even-even isobars for a given (even) value of A.

The modes of decay for isobars with A = 124 are shown in Figure 19. Neither β^+ decay of ^{124}Xe nor β^- decay of ^{124}Sn is energetically possible. ^{124}Sb decays by β^- emission and ^{124}I

[171] The $2m_e$ reflects the fact that the mass of the nucleus has decreased by m_e through the creation of a proton and the atomic number of the product nucleus has decreased by one unit, or an additional decrease of m_e in the mass of the atom as a whole.

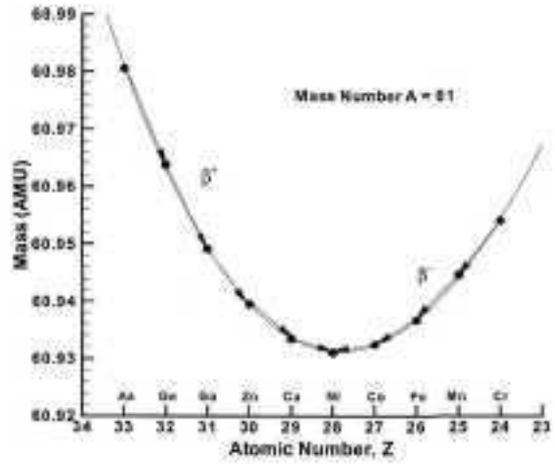


Figure 17. Masses of Isobars with A = 61. *One stable isobar, ^{61}Ni, exists.*

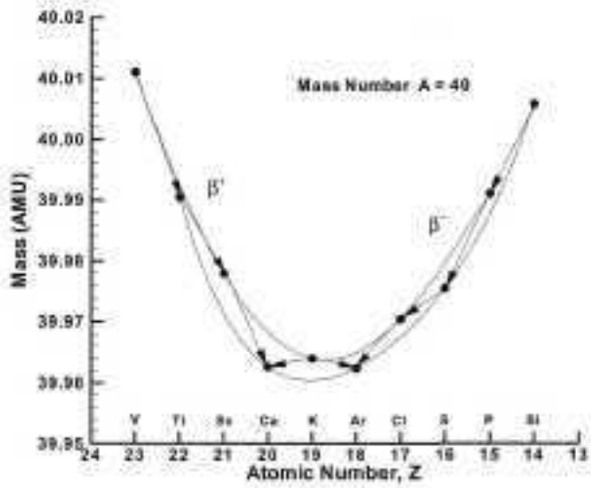


Figure 18. Masses of Isobars with A = 40. *Two stable isobars, ^{40}Ca and ^{40}Ar, exist.*

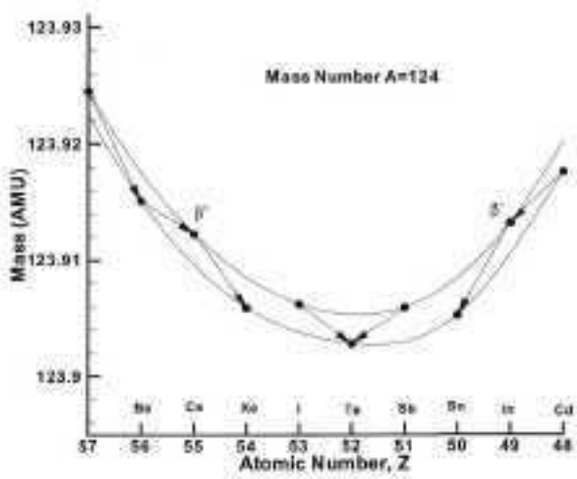


Figure 19. Masses of Isobars with A = 124. *Three stable isobars, ^{124}Xe, ^{124}Te, and ^{124}Sn, exist.*

decays by ε and β^+, both to ^{124}Te. Three stable even-even isobars, then, exist for this value of A—^{124}Xe, ^{124}Te and ^{124}Sn.

In the discussion of fission yields above, shielded nuclides were mentioned. Examples of these nuclides are ^{82}Br, ^{86}Rb, ^{100}Tc, ^{124}Sb, and ^{134}Cs; all are odd-odd nuclides. These nuclides all have a stable isobar one unit lower in Z so that they cannot be formed as a daughter product in a beta decay chain. The fission yields of these nuclides are due to their direct formation as primary fission products.

Beta decay processes are classified according to the differences in spin and parity between the initial and final states. Within a given class, the half-life is roughly proportional to the reciprocal of the 5th power of the decay energy. The fastest decays are those for which there is no change in spin or parity. Figure 20 shows the relationship between decay energies and half-lives for the 0^+ states of a series of N = Z odd-odd nuclides which decay predominantly to the 0+ ground states of even-even daughters. The solid line is $t_{1/2} \sim 7500/E^5$ ($t_{1/2}$ in seconds, E in MeV). A great majority of all beta decay half-lives falls near or above this curve.

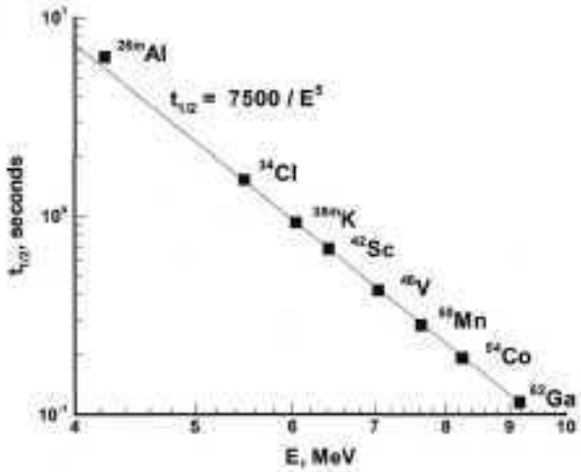


Figure 20. β^+ Decay Half-Lives of 0^+ States of Z = N Odd-Odd Nuclides Versus Decay Energy.

Alpha Decay

For nuclides with large A, alpha decay is a possible mode of decay. For alpha decay,

$$^{A}_{Z}X \rightarrow ^{A-4}_{Z-2}Y + ^{4}_{2}He$$

and for the decay to occur, the masses must satisfy the following:

$$M(A,Z) > M(A-4,Z-2) + M(4,2)$$

As mentioned above, the binding energy per nucleon for the alpha particle is about 7.1 MeV and the total binding energy is about 28.3 MeV. For some nuclides with A ≈140, the binding energy per nucleon is of the order of 7 MeV and decay by alpha emission is possible. Alpha decay becomes a dominant decay mode for proton-rich nuclides with A ≥ 160 and for neutron-rich nuclides with A ≥ 180.

Alpha decay half-lives depend strongly on the atomic number of the daughter nucleus, on the available energy, and on other factors such as whether Z and A are even or odd. For a given atomic number, the dependence of half-life on the energy of the emitted alpha particle is extremely steep. As shown in Figure 21, increasing the alpha energy from ~4 MeV to ~9 MeV for even-A uranium isotopes is associated with a decrease of about twenty orders of magnitude in half-life, compared to only about two

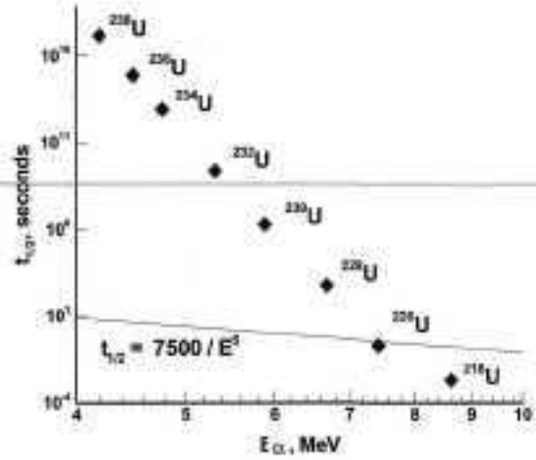


Figure 21. Alpha Decay Half-Lives of Even-A Uranium Isotopes Versus Alpha Energy.

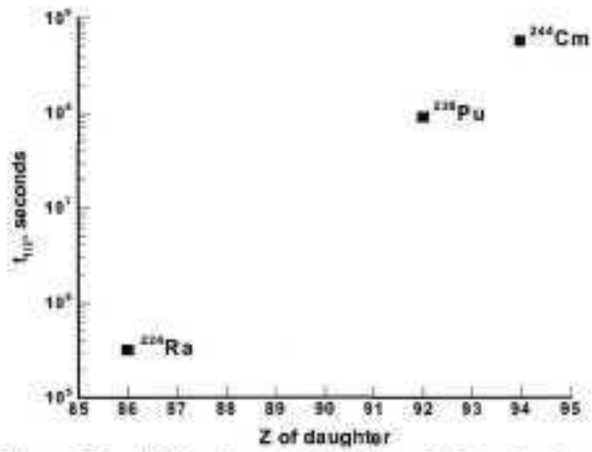


Figure 22. Alpha Decay Half-Lives of Even-A, Even-Z Nuclides With 5.68 MeV < E_α < 5.81 MeV Versus Z of the Daughter.

orders of magnitude for beta decay over a similar energy range. For a given alpha energy, the half-life increases roughly exponentially with the daughter atomic number. Figure 22 illustrates this effect for a series of even-A, even-Z alpha emitters with similar alpha energies.

Spontaneous Fission

For heavy nuclei, spontaneous fission (SF)

$$^{A}_{Z}X \rightarrow ^{A'}_{Z'}Y + ^{A-A'}_{Z-Z'}W$$

becomes possible if

$$M(A,Z) > M(A',Z') + M(A-A',Z-Z')$$

Because the maximum in the binding energy per nucleon occurs near A = 60 (see Figure 15, page 27), all nuclides with A greater than approximately 100 are unstable with respect to spontaneous fission. However, measurable rates of spontaneous fission are found only for the heaviest nuclides (A > 230) because of the high Coulomb barriers for the emission of the fission fragments. For very heavy nuclei, then, spontaneous fission becomes a major mode of decay.

Table 5. Cluster Decay Nuclides (See Reference 172)				
Nuclide	Cluster	Q(MeV)	Branching Ratio (relative to α decay)	Ref.
^{114}Ba	^{12}C	18.3-20.5	~3 E-3	173
^{221}Fr	^{14}C	31.28	(8.14 ± 1.14) E-13	174
^{221}Ra	^{14}C	32.39	(1.15 ± 0.91) E-12	174
^{222}Ra	^{14}C	33.05	(3.7 ± 0.6) E-10	175
^{223}Ra	^{14}C	31.85	(8.5 ± 2.5)E-10	193
^{224}Ra	^{14}C	30.54	(6.1 ± 1.0) E-10	175
^{225}Ac	^{14}C	30.48	(6.0 ± 1.3) E-12	176
^{226}Ra	^{14}C	28.21	(2.9 ± 1.0) E-11	177
^{228}Th	^{20}O	44.72	(1.13 ± 0.22) E-13	178
^{231}Pa	^{23}F	51.84	($9.97^{+22.9}_{-8.26}$) E-15	179
^{230}Th	^{24}Ne	57.78	(5.6 ± 1.0) E-13	180
^{232}Th	^{24,26}Ne	55.62, 55.97	<2.82 E-12	181
^{231}Pa	^{24}Ne	60.42	(1.34 ± 0.17) E-11	179
^{232}U	^{24}Ne	62.31	(2.0 ± 0.5) E-12	182
^{233}U	^{24,25}Ne	60.50, 60.75	(7.5 ± 2.5) E-13	183
^{234}U	^{24,26}Ne	58.84, 59.47	(9.06 ± 6.60) E-14	184
^{235}U	^{24,25}Ne	57.36, 57.83	(8.06 ± 4.32) E-12	185
^{236}U	^{24,26}Ne	55.96, 56.75	<9.2 E-12	185
^{232}U	^{28}Mg	74.32	<1.18 E-13	186
^{233}U	^{28}Mg	74.24	<1.30 E-15	187
^{234}U	^{28}Mg	74.13	(1.5 ± 0.5) E-13	184
^{235}U	^{28,26}Mg	72.20, 72.61	<1.8 E-12	185
^{236}U	^{28,30}Mg	71.69, 72.51	2.0 E-13	188
^{237}Np	^{30}Mg	75.02	<1.8 E-14	180
^{236}Pu	^{28}Mg	79.67	2.0 E-14	189
^{238}Pu	^{28,30}Mg	75.93, 77.03	(5.62 ± 3.97) E-17	190
^{238}Pu	^{32}Si	91.21	(1.38 ± 0.50) E-16	190
^{240}Pu	^{34}Si	90.95	<6 E-15	191
^{241}Am	^{34}Si	93.84	<2.6 E-13	192

Cluster Decay

In addition to α particle decay, certain heavy mass nuclei have been observed to decay by emitting ^{12}C, ^{14}C, ^{20}O, ^{24}Ne, ^{28}Mg, or ^{32}Si at extremely low rates. This form of decay has been designated "Cluster Radioactivity," and was first observed[193] in the emission of ^{14}C from ^{223}Ra. Since 1984, Cluster Radioactivity has been observed in 22 nuclides. Cluster Radioactivity begins with the formation of a cluster of nucleons within the nucleus, with the cluster tunneling through the potential barrier to decay.[194] Table 5 shows the branching ratios for cluster decay relative to α decay are on the order of 10^{-10} and lower and detecting this type of decay requires sensitive equipment. The table lists the nuclides observed to decay through Cluster Radioactivity, the emitted isotopes and the Q-value for the decay. Both experimental and theoretical literature can be found in Reference 172.

Drip Lines

For a given value of the mass number, A, there are limits on the possible values of the proton number, Z, and the neutron number, N, for which bound nuclei can exist. The value of Z or N for which the last nucleon is no longer bound (the binding energy is negative) and for which the nucleus decays on the time scale of 10^{-22} seconds or faster, defines the "drip line." Near the proton drip line, the Coulomb force limits the addition of protons; near the neutron drip line, the gain in binding energy with the addition of neutrons is not large enough to actually bind the neutrons. The predictions of various nuclear models as to the exact locations of the drip lines vary greatly. Experimental knowledge of the drip line locations becomes increasingly uncertain with increasing mass. Nuclear reaction experiments using high-energy beams of heavy ions give a distribution of products whose yields vary smoothly for particle-stable nuclides, but drop abruptly to zero at the drip line. For A up to ~30, yields of nuclides just inside the drip lines have often been high enough to provide a clear contrast with the zero yields beyond the drip lines. Between A~30 and A~50, the yields in recent experiments[195] for nuclides near the predicted drip lines have been barely above the detection limit (five or fewer ions). The number of nuclides that are predicted to be bound but that have not been observed at all grows rapidly above A~50. Future increases in beam intensities and detection sensitivities should establish the drip line locations up to higher masses.

For some nuclei near the proton or neutron drip lines, the weakly bound constituent protons or neutrons are in spatially extended diffuse orbitals or "halos"[196] around the core. Neutron halos have been found in ^{11}Li, ^{11}Be, ^{14}Be, ^{14}B, ^{19}C and ^{22}C. There is evidence for a proton halo in ^{8}B. The ^{11}Li nucleus appears to be a ^{9}Li core with two weakly bound neutrons forming the halo. The radius of the ^{11}Li halo is more than twice that of the core and the ^{11}Li nucleus is almost as large as the ^{208}Pb nucleus (which contains 197 more nucleons).

An Island of Stability for Superheavy Elements

The detection, by I. Curie and F. Joliot, of a short-lived radioactive isotope, ^{30}P, produced by the bombardment of aluminum with alpha particles (see page 12) showed that new isotopes could be "artificially" produced. In 1934, Fermi produced a new isotope of lead by bombarding a thallium target with slow neutrons. The capture of the neutron by the thallium nucleus, followed by β^- decay, produced the new nuclide which was one atomic number higher than the original target.

Nuclides with Z > 92 do not exist in nature. Beginning in about 1940, studies were undertaken to produce these materials using selected nuclear reactions. These studies (along with primary references) are summarized in Table 6, page 32. As described earlier, Macmillan and Abelson at the University of California at Berkeley (LBNL) first produced ^{239}Np (Z = 93) from the β^- decay of ^{239}U which had been formed by thermal neutron capture of ^{238}U. Plutonium was the next transuranium element to be discovered. At the Lawrence Berkeley Laboratory, a ^{238}U

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Table 6. Synthesis of Isotopes with Z > 92

Z	Element Name	First Isotope Synthesized					Longest Lived Isotope Known	
		Isotope	Half-Life	Date	Laboratory	Reaction	Isotope	Half-Life
93	Neptunium (Np)	^{239}Np	2.355 d	1940	LBNL[116]	$^{238}U + n \rightarrow {}^{239}U + \gamma$ $^{239}U \rightarrow {}^{239}Np + \beta^-$	^{237}Np	2.14E6 a
94	Plutonium (Pu)	^{238}Pu	87.7 a	1941	LBNL[117]	$^{238}U + {}^{2}H \rightarrow {}^{238}Np + 2n$ $^{238}Np \rightarrow {}^{238}Pu + \beta^-$	^{244}Pu	3.75E5 a
95	Americium (Am)	^{241}Am	437.7 a	1945	Metallurgical Laboratories, Univ. of Chicago[197]	$^{239}Pu + n \rightarrow {}^{240}Pu + \gamma$ $^{240}Pu + n \rightarrow {}^{241}Pu + \gamma$ $^{241}Pu \rightarrow {}^{241}Am + \beta^-$	^{243}Am	7.37E3 a
96	Curium (Cm)	^{242}Cm	162.8 d	1944	Metallurgical Laboratories, Univ. of Chicago[197]	$^{239}Pu + {}^{4}He \rightarrow {}^{242}Cm + n$	^{247}Cm	1.56E7 a
97	Berkelium (Bk)	^{243}Bk	4.5 h	1949	LBNL[198]	$^{241}Am + {}^{4}He \rightarrow {}^{243}Bk + 2n$	^{247}Bk	1.4E3 a
98	Californium (Cf)	^{245}Cf	44 m	1950	LBNL[199]	$^{242}Cm + {}^{4}He \rightarrow {}^{245}Cf + n$	^{251}Cf	9.0E2 a
99	Einsteinium (Es)	^{253}Es	20.47 d	1952	LBNL, ANL, LASL[200]	"Mike" thermonuclear explosion	^{252}Es	1.29 a
100	Fermium (Fm)	^{255}Fm	20.1 h	1954	LBNL, ANL, LASL[200]	"Mike" thermonuclear explosion	^{257}Fm	100.5 d
101	Mendelevium (Md)	^{256}Md	1.30 h	1955	LBNL[201]	$^{253}Es + {}^{4}He \rightarrow {}^{256}Md + n$	^{258}Md	51.5 d
102	Nobelium (No)	^{252}No	2.3 s	1958	JINR[202]	$^{244}Cm + {}^{12}C \rightarrow {}^{252}No + 4n$	^{254}No	55 s
103	Lawrencium (Lr)	^{258}Lr	3.9 s	1961	LBNL[203], JINR	$^{250}Cf + {}^{11}B \rightarrow {}^{258}Lr + 3n$	^{262}Lr	3.6 h
104	Rutherfordium (Rf)	^{257}Rf	4.7 s	1969	LBNL[204], JINR	$^{249}Cf + {}^{12}C \rightarrow {}^{257}Rf + 4n$	^{261}Rf	1.1 m
105	Dubnium (Db)	^{260}Db	1.5 s	1970	LBNL[205], JINR	$^{249}Cf + {}^{15}N \rightarrow {}^{260}Db + 4n$	^{262}Db	34 s
106	Seaborgium (Sg)	^{263}Sg	0.8 s	1974	LBNL[206]	$^{249}Cf + {}^{18}O \rightarrow {}^{263}Sg + 4n$	^{266}Sg	21 s
107	Bohrium (Bh)	^{262}Bh	0.10 s	1981	GSI[207]	$^{209}Bi + {}^{54}Cr \rightarrow {}^{262}Bh + n$	^{267}Bh	17 s
108	Hassium (Hs)	^{265}Hs	1.6 ms	1984	GSI[208]	$^{208}Pb + {}^{58}Fe \rightarrow {}^{265}Hs + n$	^{277}Hs	16.5 s
109	Meitnerium (Mt)	^{266}Mt	3.4 ms	1982	GSI[209]	$^{209}Bi + {}^{58}Fe \rightarrow {}^{266}Mt + n$	^{268}Mt	0.07 s
110	Darmstadtium (Ds)	^{269}Ds	270^{+1300}_{-120} µs	1995	GSI[210], JINR	$^{208}Pb + {}^{62}Ni \rightarrow {}^{269}Ds + n$	^{281}Ds	1.6 m
111	—	^{272}111	$1.5^{+2.0}_{-0.5}$ ms	1995	GSI[211], JINR	$^{209}Bi + {}^{64}Ni \rightarrow {}^{272}111 + n$	^{272}111	1.5 ms
112	—	^{277}112	240^{+430}_{-90} µs	1996	GSI[212], JINR	$^{208}Pb + {}^{70}Zn \rightarrow {}^{277}112 + n$	^{285}112	15.4 m
113								
114	—	^{292}114	30.4 s	1999	JINR[213], LLNL	$^{244}Pu + {}^{48}Ca \rightarrow {}^{292}114$	^{292}114	30.4 s
115								
116	—	^{292}116	46.9 ms	2000	JINR[214]	$^{248}Cm + {}^{48}Ca \rightarrow {}^{292}116 + 4n$	^{292}116	46.9 ms

target was bombarded by deuterons to produce ^{238}Np, which subsequently underwent β^- decay to form ^{238}Pu (Z = 94):

$$^{238}_{92}U + {}^{2}_{1}H \rightarrow {}^{238}_{93}Np + 2{}^{1}_{0}n$$

and

$$^{238}_{93}Np \rightarrow {}^{238}_{94}Pu + \beta^- + \bar{\nu}$$

Studies of the debris from the "Mike" thermonuclear explosion carried out in the Pacific in 1952 led to the discovery of two new elements. In a collaborative study carried out by scientists from Argonne National Laboratory (ANL), Los Alamos National Laboratory (LANL) and the University of California at Berkeley, the isotopes ^{253}Es (Z = 99) and ^{255}Fm (Z = 100) were found in the debris. Beginning with ^{238}U, these nuclides were formed in the explosion through a combination of many neutron capture reactions, made possible because of the extremely high neutron flux, and β^- decay of some of the resulting products.

Some isotopes in the Z = 93 to 99 region can be produced in weighable quantities. To synthesize the new elements up to Z = 101, a "stepping-stone" procedure was used. Reactions of neutrons, deuterons, or alpha particles incident on newly synthesized high-Z materials produced new nuclides with still higher Z values. By the mid-1950's elements up to Z = 101 had been produced (see Seaborg and Loveland[118] for details).

Producing weighable quantities of isotopes with Z>100 is not possible due to the short alpha decay and fission half-lives of these heavier elements, so light particle induced reactions cannot be used to produce new, higher Z materials. To produce elements with Z>101, heavy ion accelerators were required. Between 1958 and 1974, at Berkeley and the Joint Institute for Nuclear Research (JINR) at Dubna, Russia, nuclides with $102 \leq Z \leq 106$ were produced with beams of carbon, boron, nitrogen or oxygen ions bombarding californium and curium targets.

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Until about 1970 it was thought that a limit to the periodic table would be reached at about element 108. Beyond this, it was reasoned, the half-lives of the heavier materials, particularly for decay by spontaneous fission, would probably be so short as to prevent them from being observed.

As mentioned earlier, shell model calculations successfully explained the relative stability of those nuclides having "magic" proton or neutron numbers 2, 8, 20, 50 and 82 and, for neutrons, 126. The calculations also predicted that the next magic number for protons would be 126 and for neutrons, 184. A number of nuclear structure calculations carried out in the late 1960's and early 1970's, however, suggested both Z = 114 and 126 as possible closed subshells, with more recent work[215] supporting a closed subshell at Z = 114. Nuclei with 114 protons and 184 neutrons were predicted to have filled both neutron and proton shells (doubly magic) and would be particularly stable. Some early calculations indicted that these superheavy nuclides, having "shell stabilization," could have half-lives of the order of the age[216] of the Universe. In this mass region, then, nuclides with relatively long half-lives should exist, forming an "island of stability" rising out of the "sea of instability." More recent calculations, while still supporting the existence of the island of stability, predict somewhat shorter half-lives.

The reactions that had been used to produce the elements above Z = 92 and below Z = 107 used heavy targets, $92 \leq Z \leq 98$, and "light" incident particles (see Table 6, page 32). The use of these types of reactions to produce elements with $Z \geq 107$ was not successful. For the "fusion" of the target and projectile to occur, enough energy must be available to overcome the Coulomb repulsion between the two ions. The resulting new compound system will be highly excited, however, and will likely undergo fission to reach a more stable configuration.

To synthesize the heavier elements, scientists at the Institute of Heavy Ion Research (GSI) in Darmstadt, Germany and the JINR at Dubna, Russia used heavier incident particles to bombard somewhat lighter targets. This results in lower excitation energy in the compound system, decreasing the tendency to undergo fission. If the excitation energy is low enough, most of the excess energy is lost by the emission of neutrons from the nucleus—not by fission.

Both the production and detection of the high Z elements are extremely difficult. The low cross sections for the reactions involved have led to the development of ion sources and accelerators capable of delivering high currents of heavy ions.[217] Efficient means of isolating and detecting the products of these reactions are also necessary.[218] The identification of a new nuclide is carried out by correlating its decay with the decays of daughter products whose emissions, energies and half-lives are known from previous measurements. Bohrium (Z = 107) was first produced at GSI using the reaction[207]

$$^{209}Bi + {}^{54}Cr \rightarrow {}^{262}Bh^{*} \rightarrow {}^{262}Bh + n$$

Five decay chains of ^{262}Bh were observed, one of which is shown in Figure 23. Specific decays in each of the decay chains were correlated with known decays in either ^{254}Lr, ^{250}Fm, or ^{250}Md. The original parent was, then, assigned with a high probability to be ^{262}Bh.

Subsequent work at GSI, JINR and LLNL has produced isotopes of elements with Z = 108, 109, 110, 111, 112, 114 and 116. The reactions used to produce these new materials are briefly summarized in Table 6 (page 32). Most of the superheavy nuclides synthesized thus far decay by alpha emission—the probability of fission is greatly reduced by the shell stabilization effects mentioned above. The production of isotopes with Z = 114 by JINR and LLNL used a neutron-rich projectile, ^{48}Ca, on

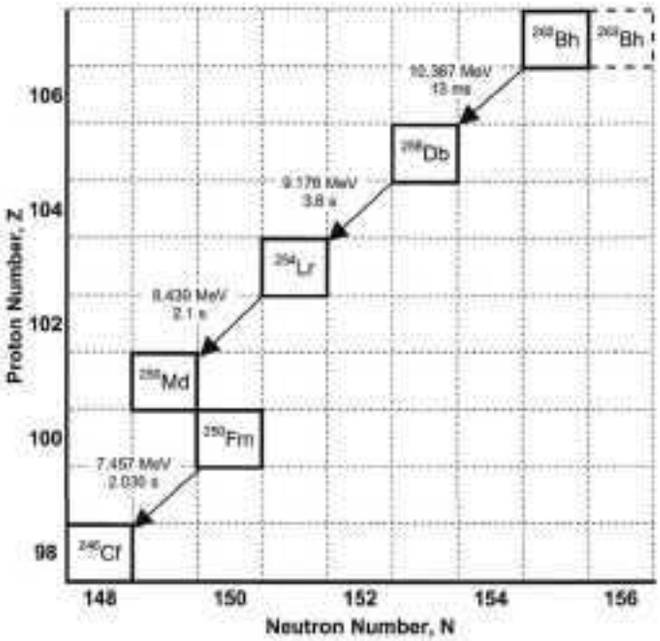

Figure 23. One of the Decay Chains Observed in the Synthesis of ^{262}Bh (from Reference 207). *Alpha decay energies and observed half-lives are shown for each decay.*

targets of both ^{244}Pu and ^{242}Pu, giving the desired value of Z.[213] The reaction products were closer to the "island of stability" in terms of the number of neutrons, N.

The use of the ^{48}Ca reaction to reach Z = 110, 112, and 114 has been well documented and the results reproduced in the literature.[210] The element Z = 116 was independently identified[214] in 2000 using the ^{248}Cm(^{48}Ca,xn)^{292}116 reaction at JINR. The identification of "one" decay to ^{288}114 (and subsequent known daughter decays) with a half-life of 46.9 ms was used for the identification (see Figure 24).

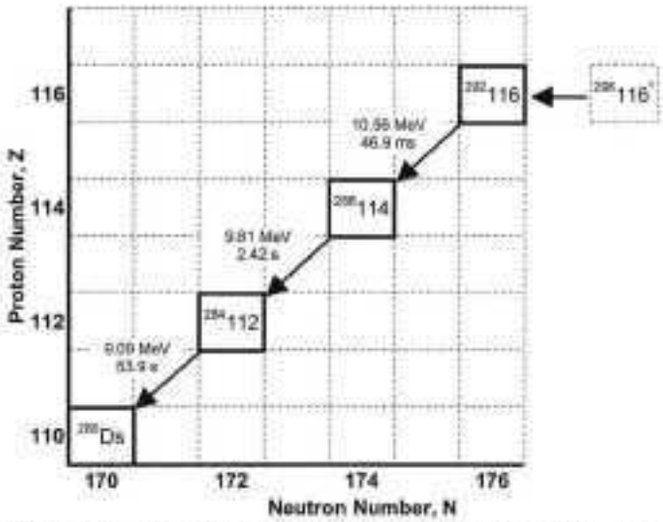

Figure 24. The Reported Decay Chain for Z = 116. *The decays from Z = 114 had previously been identified (from Reference 213). Alpha decay energies and observed half-lives are shown for each decay.*

[207] see G. Hermann, Science **280** 543 (1998).
[210] L. Knox, N. Christensen and C. Skordis, Astrophys. J. **563** L95 (2001).
[213] P. Armbruster and F.P. Hessberger, Scientific American, Sept. 1998, pp. 72-77.
[214] S. Hofmann and G. Munzenberg, Rev. Mod. Phys. **72** 733 (2000).
[215] Y.T. Oganessian et al., Acta Phys. Slov. **49** 65 (1999).

Related Topics

Origin of the Naturally-Occurring Nuclides

A "Big Bang" model of the history of the universe[220] has become widely, but not universally, accepted.[221] According to the Big Bang model, the universe started out unimaginably small, dense and hot, about 14 billion years ago, and has been expanding and cooling (on the average) ever since.[218] George Gamow and others initially proposed that all the elements could have been created during the Big Bang.[222, 223] Neutrons and protons would begin to condense into nuclei when the temperature of the expanding universe dropped below about 10^9 kelvin. However, it was soon realized that the building up of heavy nuclei during the Big Bang could not have continued very far, because collisions between nuclei became less frequent as the universe cooled, and the thermal energy of the nuclei became too low to overcome the electrostatic repulsion of their positive charges. The absence of stable nuclides with mass 5 or 8 hindered the buildup of the heavier nuclides.

As an alternative source of heavy elements, Burbidge, Burbidge, Fowler and Hoyle, (B^2FH), proposed in 1957 a set of mechanisms by which most of the elements were formed in stars and a few by cosmic ray processes.[224] Subsequent observations, laboratory experiments, and calculational modeling have supported many of their hypotheses and filled in some of the details.[225] One difficulty in comparing theory with observation is that stars have enormous differences in composition, density, and temperature between their inner and outer regions, but the light we see comes only from the outer surface. Stellar spectra correspond to effective temperatures from ~2×10^3 to ~4×10^3 K, while fusion reactions begin at ~2×10^7 K and core temperatures may eventually approach 1×10^{10} K. Although the temperature differences between core and surface are large, the distances are also large, so thermal gradients are rarely large enough to drive convective mixing of matter from inner regions to the surface. Another difficulty is that several of the reaction sequences go through unstable nuclides that are difficult to study in the laboratory. Also, reactions that are too slow to measure with precision in the laboratory can be important in stars, where high density compensates for low cross sections.

One weakness in the "B^2FH" model became apparent when more data on the current cosmic abundances of light nuclides became available. The abundance of deuterium, in particular, is too high to be explained by stellar or cosmic ray processes. Deuterium is consumed more easily than it is produced, and, if cosmic rays were the source of deuterium, they would also have produced much more than the observed amount of ^{7}Li.

Most astrophysicists now attribute the formation of the lightest nuclides to the Big Bang and the heavier ones to subsequent stellar and cosmic ray processes. Modern predictions of the relative abundance of the nuclides formed in the Big Bang are highly dependent on the density of the universe as it cooled through the temperature range where fusion reactions were possible. The composition of the early universe can be inferred from analysis of stellar spectra, cosmic rays, and meteorites, with corrections to account for changes believed to have occurred over time. The resulting estimates of primordial abundance for the five light nuclides ^{1}H, ^{2}H, ^{3}He, ^{4}He, and ^{7}Li are all consistent with about the same value of the density of the universe. This consistency is considered to be strong evidence in favor of the Big Bang model. This density is low compared to the densities in stars, where fusion of lighter nuclides into heavier ones is occurring today.

The energy necessary to overcome the coulomb barrier between higher-Z nuclei comes from the release of gravitational potential energy as the star contracts. More massive stars are able to achieve higher temperatures and greater densities and do so on a shorter time scale than less massive stars. Only a small subset of stars, the most massive ones, are believed to have produced most of the nuclides heavier than ^{7}Li. Much of the heavier matter produced will never escape to be dispersed through the rest of the universe.

Fusion of hydrogen into helium is the first source of nuclear energy in the life cycle of a star. The subsequent steps depend strongly on the star's mass. The Sun and stars of comparable or lower mass remain in the hydrogen-burning stage for billions of years. Stars of the Sun's mass have contributed little to the production of heavier elements since the Big Bang. More massive stars consume their hydrogen more rapidly. Further gravitational contraction then heats them to the temperatures required to fuse helium into heavier nuclides. If the star is massive enough, it evolves through stages of carbon burning, neon burning, oxygen burning, and silicon burning. Later stages take less time than earlier stages and release less energy per unit mass of fuel, providing less resistance to the gravitational contraction.

The details of the earlier stages depend greatly on the relative rates of individual nuclear reactions, but in the later stages when temperature and density become sufficiently high, a quasi-equilibrium state can be reached, where rates are less important than binding energies. Nuclei near the peak of the binding energy curve dominate the composition in the quasi-equilibrium state. Evidence for quasi-equilibrium comes from the high abundance of ^{56}Fe and nearby nuclides, and from the observation of supernova remnants. The prodigious light output of the remnants is powered by radioactive decay of proton-rich nuclides near the binding energy peak, such as ^{56}Ni, ^{56}Co, ^{57}Ni, and ^{44}Ti. This conclusion is supported both by the rate of decay and by the observation of characteristic gamma-ray lines. (Heat and light are produced when some of the nuclear radiation is absorbed by the matter that was ejected in the explosion.)

Hydrogen burning begins with the cycle:

$$p + p \rightarrow {}^2H + \beta^+$$
$${}^2H + p \rightarrow {}^3He + \gamma$$
$${}^3He + {}^3He \rightarrow {}^4He + 2p$$

The first reaction is slow compared to the next two, so the intermediate products, ^{2}H and ^{3}He, remain in low concentrations. At higher temperatures, ^{7}Be is formed by the reaction of ^{3}He with ^{4}He. Proton reactions on ^{7}Be and the decay product, ^{7}Li, provide additional channels for the formation of ^{4}He.

Helium burning begins with the near-simultaneous fusion of three ^{4}He nuclei to form ^{12}C. The third ^{4}He can add to the unstable ^{8}Be intermediate before it decays only because the reaction can pass through an excited state of ^{12}C in the necessary energy range. (This state was postulated to exist in order to explain the astronomical observations before it was discovered in laboratory experiments.)[226] Helium can fuse with ^{12}C to form ^{16}O,

[220] M.J. Rees, Science **290** 1919 (2000).

[221] F. Hoyle, G. Burbidge, and J.V. Narlikar, *A Different Approach to Cosmology: From a Static Universe Through the Big Bang Towards Reality*, (Cambridge University Press, New York, 2000).

[222] G. Gamow, Phys. Rev. **70** 572 (1946).

[223] R.A. Alpher, H. Bethe, and G. Gamow, Phys. Rev. **73** 803 (1948); R.A. Alpher and R.C. Herman, Phys. Rev. **75** 1089 (1949).

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[225] G. Wallerstein *et al.*, Rev. Mod. Phys. **69** 995 (1997); A. Burrows, Nature **403** 727 (2000); D. Arnett, *Supernovae and Nucleosynthesis: An Introduction to the History of Matter from the Big Bang to the Present*, (Princeton University Press, 1996); C.J. Copi, D.N. Schramm and M.S. Turner, Science **267** 192 (1995); D.N. Schramm and M.S. Turner, Rev. Mod. Phys. **70** 303 (1998); Big Bang Nucleosynthesis, http://cfa-www.harvard.edu/~mwhite/darkmatter/bbn.html; S. Sneddon and J.J. Coran, Science **229** 70 (2003).

[226] F. Hoyle, D.N.F. Dunbar, W.A. Wenzel and W. Whaling, Phys. Rev. **92** 1095 (1953).

but early predictions that ^{20}Ne would be produced by fusion of helium with ^{16}O have not been supported by later studies. ^{20}Ne lacks states of the necessary spin and parity in the appropriate energy range for this reaction to be important in stars. Instead, neon is formed by fusion of pairs of carbon nuclei and by successive proton capture reactions on the lighter elements.

When hydrogen and carbon are both present, hydrogen burning can be catalyzed by carbon in the CN cycle:

$$^{12}C(p,\gamma)^{13}N$$
$$^{13}N \rightarrow \beta^{+} + {}^{13}C$$
$$^{13}C(p,\gamma)^{14}N$$
$$^{14}N(p,\gamma)^{15}O$$
$$^{15}O \rightarrow \beta^{+} + {}^{15}N$$
$$^{15}N(p,\alpha)^{12}C$$

This cycle can be written in a more convenient form as:

$$^{12}C(p,\gamma)^{13}N(\beta^{+})^{13}C(p,\gamma)^{14}N(p,\gamma)^{15}O(\beta^{+})^{15}N(p,\alpha)^{12}C.$$

A small fraction of the ^{15}N is converted to ^{16}O by a (p,γ) reaction, but is returned to the CN cycle by the NO cycle:

$$^{15}O(p,\gamma)^{17}F(\beta^{+})^{17}O(p,\alpha)^{14}N$$

The CN and NO cycles require higher temperatures than the p + p reaction but at sufficiently high temperature, they proceed at a much greater rate. At still higher temperatures, the $^{13}N(p,\gamma)^{14}O$ reaction competes with the beta decay of ^{13}N, initiating a "hot CNO cycle" that leaves as a signature a $^{15}N/^{14}N$ ratio 25,000 times higher than the slower beta-limited cycle. The hot CNO cycle is believed to be the main source of energy in nova explosions.

As hydrogen is depleted and temperature increases, pairs of carbon nuclei begin to fuse into isotopes of neon and magnesium, with emission of alpha particles, protons, and neutrons. As the temperature increases further, the gamma-ray flux becomes sufficient to photodissociate neon.

$$^{20}Ne(\gamma,\alpha)^{16}O$$

The released alpha particles then fuse with other neon nuclei:

$$^{20}Ne(\alpha,\gamma)^{24}Mg$$

As the temperature increases still further, pairs of ^{16}O nuclei begin to fuse, forming isotopes of silicon, phosphorus and sulfur, with the emission of alphas and lighter particles. At still higher temperatures (3-3.5 x 10^{9} K), more photodissociation reactions become possible. Alphas, protons, and neutrons released by photodissociation then fuse with silicon and other nuclei, forming products up to the binding energy peak.

Four more chain processes are required to explain the abundances of the nuclides beyond the peak of the binding energy curve. One of these is the "s-process", in which a slow succession of alternating neutron captures and β^{-} decays creates nuclides lying along the valley of stability. The main sources of neutrons for the s-process are the $^{13}C(\alpha,n)^{16}O$ and $^{22}Ne(\alpha,n)^{25}Mg$ reactions. Stable nuclides with magic neutron numbers and low neutron cross sections act as bottlenecks in the s-process and are produced in the highest yield. The s-process cannot produce elements beyond bismuth (Z = 83), because they decay by α-emission back to lighter elements before they can absorb the next neutron.

Under conditions of much higher neutron density, successive neutron captures take place faster than the products can decay. The rapid neutron capture, or r-process, creates nuclides near the neutron drip line all the way up to the actinides. While the r-process is occurring, it also preferentially produces nuclei with magic neutron numbers, but with fewer protons than those preferentially produced by the s-process. Subsequent β^{-} decay converts these drip line nuclei to stable nuclei with neutron numbers less than the magic neutron numbers, and mass numbers less than the nuclei produced in maximum yield in the s-process.

Some of the nuclides on the proton-rich side of the valley of stability, such as ^{74}Se and ^{92}Mo, cannot be produced by neutron capture. These nuclides are believed to be produced by proton capture and (γ,n) reactions, which may be slow (p-process) or fast (rp-process) relative to beta decay.

The conditions necessary for the r-process are believed to exist during the explosion of a type II supernova, when the core of a massive star collapses to a neutron star or a black hole. The rp-process is likely to occur in a more hydrogen-rich environment, such as occurs in a nova explosion following transfer of hydrogen from a nearby star onto a star that has already evolved beyond hydrogen burning.

Some of the lightest nuclides, ^{6}Li, ^{9}Be, ^{10}B, and ^{11}B, are bypassed by the stellar nucleosynthesis chains, and would be quickly consumed in stars by reactions with protons, neutrons, and alpha particles. Their abundances are explained in part by fragmentation of carbon and oxygen nuclei in interstellar cosmic ray collisions.

A more thorough treatment of nucleosynthesis and stellar evolution can be found in Reference 225 and in the Additional Reading Material section on page 37.

Dark Matter and the Solar Neutrino Problem

Two of the most active areas of current research at the overlap of nuclear physics and astronomy are the "dark matter problem"[227] and the "solar neutrino problem".[228] Several lines of evidence (the rotational speeds of galaxies, for example) indicate that the universe contains more mass than the sum of the mass of all the luminous stars. Some of the "dark matter" must simply be ordinary atoms and molecules that are not hot enough to emit light. However, if all the mass of dark matter were ordinary matter (mostly protons and neutrons by weight) it would be inconsistent with the density of ordinary matter inferred through the Big Bang model from the relative abundances of light nuclides. The search continues for ordinary matter, which could reside in the outer regions of galaxies in MACHOs (massive astrophysical compact halo objects), and for more exotic forms of matter such as WIMPs (weakly-interacting massive particles).

The solar neutrino problem also highlights some weaknesses in current understanding. In contrast to photons, neutrinos can escape from the very center of a star to reach the earth. Since the 1970's, large-scale experiments have sought to measure the flux from the sun of the neutrinos of particularly high energy produced when ^{8}B decays. These experiments have consistently found fewer neutrino interactions in their detectors than predicted. Newer experiments are also finding fewer interactions of lower-energy neutrinos than predicted. The deficit could result from a shortcoming in the solar models. However, other large-scale experiments are beginning to show that it may instead be due to previously undetected properties of neutrinos, which are associated with a very small, but non-zero mass.[229]

Counting Atoms Not Decays

There are many applications for a technique that would count atoms of long-lived radioactive nuclides, rather than waiting for their decay, or measuring their decay rate. For example, the ^{14}C dating method[230] depends upon the measurement of the beta-counting rate of contemporary carbon of biological origin. One gram of contemporary carbon has a beta-disintegration rate of about 13.5 decays per minute per gram. However, this low decay

[227] R. Cowen, Sci. News **160** 234 (2001); B.D. Fields, Science **294** 529 (2001).
[228] R.S. Raghavan, Science **267** 45 (1995).
[229] SNO Collaboration, Phys. Rev. Lett. **87** 071301 (2001); The Neutrino Industry website: http://www.hep.anl.gov/ndk/hypertext/nuindustry.html
[230] Carbon-14 (half-life 5715 years).

rate is in the presence of ~6.5E10 atoms of ^{14}C in a one gram sample. Clearly, if ^{14}C atoms could be counted efficiently by some technique, it would be possible to determine the ^{14}C content of very small quantities of carbon. This has now been accomplished for milligram carbon samples even though the ratio of radioactive ^{14}C to stable ^{12}C was ~1E-12. The technique is ultrasensitive mass spectrometry using accelerators.

An accelerator, such as a cyclotron, can act as an extremely sensitive mass spectrometer. Only particles in the ion source with the proper charge-to-mass ratio (given by the cyclotron resonance equation) will be accelerated. Most radioactive atoms to be detected cannot be distinguished from stable contaminants of the same charge-to-mass ratio present in much higher amounts. Use of an accelerator to boost the particles to high energies makes the separation possible by measuring their different energy loss or ionization rates in a detector. For example, ^{14}C of +4 charge can be distinguished from ^{14}N of the same charge. In effect, the accelerator acts as a high-energy mass spectrometer characterized by both high mass resolution and extremely low background count rates. The sample count rate has to be low enough to avoid saturation and radiation damage effects in the detector and the sample has to be small enough to process through the cyclotron in a reasonable time. For application to the radiocarbon dating of archaeological samples, the cyclotron is tuned to accelerate the carbon isotope of interest, and the sample is introduced into the ion source, preferably as a gas. The accelerated ions with the proper signature in the energy loss detector are then counted. In this way, the sensitivity is so enhanced over counting decays that much smaller (milligram) samples can be used to obtain more precise results. The upper range of "ages" that can be dated is now potentially 40,000 to 100,000 years, rather than ~10,000 years.

The technique is not limited to radiocarbon dating. Any long-lived nuclide can potentially be detected in this way, including H^3 (tritium), ^{10}Be, ^{26}Al, ^{53}Mn, ^{36}Cl, ^{39}Ar, ^{81}Kr, ^{129}I, and even ^{205}Pb. This new technique has implications, then, for archaeology, geology, nuclear waste disposal, pollution monitoring, and measurement of nuclear properties such as half-lives. It is potentially more valuable, the longer the half-life of the radionuclide involved (and the longer it would take to observe decays rather than atoms). As examples, tritium has been used to tell how long it would take seepage to refill underground reservoirs, by measuring the age of the water. ^{36}Cl could be used to detect leakage into groundwater near nuclear power plants and waste disposal sites. ^{10}Be has been used in studies of both sea-floor spreading and manganese nodule formation. Its potential utilization with the ultrasensitive mass spectrometer technique should be practical for ages of 5 to 35 million years, or more, depending upon the size of the rock samples analyzed. Finally a more precise value of the half-life of ^{32}Si has been measured by two independent groups using this new technique. Readers interested in this technique are referred to review articles by A.E. Litherland.[231]

Radioisotopic Power Data

Radioisotopes have been used to power earth orbit satellites as well as experiments left after lunar landings. The kinetic energy of the decay products is transformed into heat energy that is converted to electric power by several techniques. Thermocouples are one conversion device, while thermionic converters are another.

The fuel for a radioisotopic power generator must be safe, reliable and low in both weight and cost. These requirements result in a rejection of radioisotopes with half-lives less than 100 days, because they would have to be replaced too often. Half-lives greater than 100 years would be rejected because the activity would be too low to obtain an initial power density greater than 0.1 watts/gram. Some of the candidates that would qualify as fuel sources for radioisotopic power generation and their calculated initial power densities in watts/gram are given in Table 7.

Table 7. Radioisotopic Power Data[232] in Watts per Gram

Nuclide	Half-Life	W/g
^{3}H	12.32 a	0.325
^{60}Co	5.271 a	17.45
^{85}Kr	10.76 a	0.590
^{90}Sr	28.78 a	0.916
^{106}Ru	1.020 a	31.8
^{137}Cs	30.07 a	0.427
^{144}Ce	284.6 d	25.5
^{147}Pm	2.6234 a	0.340
^{170}Tm	128.6 d	11.86
^{210}Po	138.38 d	141.3
^{238}Pu	87.7 a	0.558
^{241}Am	432.7 a	0.113
^{242}Cm	162.8 d	120.0
^{244}Cm	18.1 a	2.78

Pre-Fermi (Natural) Nuclear Reactor

In 1956, P.K. Kuroda[233] speculated that an assemblage of uranium and water could easily become a self-supporting chain reacting system in the early history of the Earth. These speculations were borne out by the French discovery in Oklo in the Gabon Republic in Africa of samples of natural uranium, which have been found to be depleted in ^{235}U. (There is measurably less ^{235}U in these samples than in "ordinary" natural uranium.) Samples that were slightly but definitely enriched in ^{235}U were also found. The abundances of various nuclides found in the samples did not agree with the natural abundance of these nuclides. The distribution of isotopes of particular elements, after correction for neutron capture and natural background, agreed with the fission yield distribution from ^{235}U.

From the Chart, it can be noted that neodymium has many stable isotopes at the end of short fission product chains which have large fission yields. The analysis of neodymium provided much of the information about the Oklo samples.

From radioactive dating methods, the samples are determined to be about 1.7 billion years old. The ^{235}U abundance in this natural reactor would have been 3% or more (prior to the ^{235}U decay over the past 1.7 billion years to its present 0.72% abundance). The geological evidence implies that large quantities of water were present to provide moderation of the fission neutrons. After fissioning, the samples would have been depleted in ^{235}U. Enriched samples would have resulted from the ^{238}U transforming into ^{239}Pu by neutron capture and decay and subsequent alpha decay of ^{239}Pu into ^{235}U. This conversion reaction would also have fed the reactor. Studies of the Oklo natural reactor[234] indicated that about two-thirds of the total fission events were caused by the ^{235}U initially present in the ore and nearly one-third from ^{235}U converted from ^{238}U by neutron capture and subsequent decay.

By studying the amounts of the various nuclides in the ore, a wealth of information has been obtained. This information includes values of neutron flux, time-integrated neutron flux, conversion ratio and fraction of epithermal neutrons. Finally, a determination has been made of the number of fissions that had occurred within the reactor and the percentage that had been caused by each of the fissioning nuclides ^{235}U, ^{238}U, and ^{239}Pu. It has been estimated that a total released energy of 15,000

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megawatt years was generated by the Oklo reactor. This is approximately the same energy as generated by one large modern reactor in four years. It is interesting to note that this naturally-occurring enriched converter reactor operated nearly two billion years before Fermi initiated the first man-made neutron chain reaction in 1942 with natural uranium fuel. Clearly, in this instance, man was the unknowing imitator of nature.

At least one other feature of the Oklo natural reactor has stimulated interest. This feature has been the remarkable stability of the Oklo ore deposit. The ^{235}U and ^{238}U fuels have been confined within the grains of ore. Since almost half of the residual ^{235}U in the reactor zones is the daughter product of ^{239}Pu, this means that the plutonium has been "locked-up" for a time comparable to its 2.410E4 year half-life. Also, at least one-half of thirty or so fission product elements have remained immobilized in the ore. These findings have produced much interest from scientists working on the complex problem of waste storage and disposal of radioactive fuel and wastes.

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Conversion Factors

The International System of Units, developed and maintained by the General Conference on Weights and Measures (abbreviated CGPM from the official French name) is intended as a basis for worldwide standardization of measurement units. The name International System of Units and the international abbreviation SI (from the French name Le System Internationale d'Unites) were adopted by the 11th CGPM in 1960. SI is based on a selected set of dimensionally independent SI Base Units. These are the meter, kilogram, second, ampere, kelvin, candela, and mole. In the SI system there is one and only one SI unit for each physical quantity. This is either the appropriate SI Base Unit itself or the appropriate SI Derived Unit formed by multiplication and/or division of two or more SI Base Units. For example, newtons, pascals, and joules are the names of SI Derived Units for force, pressure (stress), and energy.

Prefixes and symbols are used to form names of decimal multiples of the SI units. These are listed in the table below.

SI Prefixes		
Prefix	**Symbol**	**Multiplication Factor**
yotta	Y	10^{24}
zetta	Z	10^{21}
exa	E	10^{18}
peta	P	10^{15}
tera	T	10^{12}
giga	G	10^{9}
mega	M	10^{6}
kilo	k	10^{3}
hecto	h	10^{2}
deka	da	10^{1}
deci	d	10^{-1}
centi	c	10^{-2}
milli	m	10^{-3}
micro	μ	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}
femto	f	10^{-15}
atto	a	10^{-18}
zepto	z	10^{-21}
yocto	y	10^{-24}

In the following sections, SI units are emphasized. Each section (Length, Area, Angle, etc.) consists of two parts: one in which the conversion of other units to the appropriate SI units are given and one in which conversion factors are given for non-SI units. The conventions used are as follows: Conversion factors set in **boldface** are exact. When four or more digits are given, they are broken up into groups of three for ease of readability. Exponential nomenclature is used, i.e., 2.011 684 E+01 means $2.011\ 684 \times 10^{1}$.

The reader is cautioned that the SI system is evolving and that there will be changes from time to time as this occurs. For example, the SI Base Unit for length, the meter, was redefined by CGPM in 1983 as the length of the path traveled by light in a vacuum during a time interval of 1/299 792 458 of a second. For a more complete discussion of SI Units, the reader is referred to "Standard for Use of the International System of Units (SI): The Modern Metric System", SI10 IEEE/ASTM SI-10 1997.[235] Additional information on conversion factors can be found in the NIST Handbook 44.[236]

[235] Published by The Institute of Electrical and Electronics Engineers, Inc., 345 East 47th Street, New York, NY, 10017-2394 and the American Society for Testing and Materials (ASTM), 100 Bar Harbor Drive, West Conshohocken, PA 19428-2959, USA. (see (http://www.astm.org).

[236] See Appendix C. General Tables of Units of Measurement, NIST Handbook 44—2001 Edition, "Specifications, Tolerances, And Other Technical Requirements for Weighing and Measuring Devices" (see http://ts.nist.gov/ts/htdocs/230/235/appxc/appxc.htm).

Angle

The SI unit of plane angle is the radian (rad).[237] The radian is the plane angle between two radii of a circle, which cut off on the circumference an arc equal in length to the radius.

The SI unit of solid angle is the steradian (sr). The steradian is the solid angle which, having its vertex in the center of a sphere, cuts off an area of the surface of the sphere equal to that of a square with sides of length equal to the radius of the sphere.

Conversion to radians and steradians from other units

To convert from	to	Multiply by
degree (angle)	radian (rad)	1.745 329 E-02
grade	radian (rad)	1.570 796 E-02
minute (angle)	radian (rad)	2.908 882 E-04
second (angle)	radian (rad)	4.848 137 E-06
solid angle	steradians (sr)	1.256 637 E+01
square degree	steradians (sr)	3.046 173 E-04

Other Unit Conversions of Angle

1 degree	=	**60** minutes
1 grade (gon)	=	**9.0 E-01** degrees
1 minute	=	**60** seconds
1 mil (angle)	=	**360/6400** degrees
	=	**5.625 E-02** degrees
1 radian	=	**360°/2π**
	=	5.729 578 E+01 degrees
1 sec slope	=	4.848 137 E+00 microinches/inch
	=	0.001 inches/17.188 73 feet
1 solid angle	=	**4π** steradians
	=	1.256 637 E+01 steradian
	=	**1** sphere
1 steradian	=	7.957 747 E-02 solid angle

[237] This may be confused with the unit of absorbed dose, the rad. The absorbed dose rad should be abbreviated rd.

Length

The SI unit for length is the meter, symbol m. The meter currently in use is the length of the path traveled by light in a vacuum during a time interval of 1/299 792 458 of a second.

Since 1893 the definition of the U.S. yard has been derived from the meter. In 1959 the definition of the yard was modified to bring the U.S. yard and the yard used in other countries into agreement. The U.S. yard was changed from 3600/3937 m to 0.9144 m exactly.

In 1959 it also was decided that any data in feet derived from and published as a result of geodetic surveys within the U.S. would remain with the old standard (1 foot=1200/3937 m) until further decision. This foot is named the U.S. survey foot. Conversions to meters for U.S. survey units are given in a separate table, below.

1 foot (international) = **0.3048** m
1 foot (U.S. survey) = **1200/3937** m = 3.048 006 E-01 m

Conversion to meters from other units

To convert from	to	Multiply by	
angstrom	meter (m)	**1.0**	**E-10**
astronomical unit (international)	meter (m)	1.495 979 E+11	
foot	meter (m)	**3.048**	**E-01**
inch	meter (m)	**2.54**	**E-02**
light year	meter (m)	9.460 528 E+15	
mil (0.001 in)	meter (m)	**2.54**	**E-05**
mile, international (5280 ft)	meter (m)	**1.609 344**	**E+03**
mile (international nautical)[238]	meter (m)	**1.852**	**E+03**
parsec	meter (m)	3.085 678 E+16	
yard	meter (m)	**9.144**	**E-01**

Conversion to meters from U.S. survey[239] units

To convert from	to	Multiply by
link	meter (m)	2.011 684 E-01
survey foot	meter (m)	3.048 006 E-01
rod (16 ½ survey feet)	meter (m)	5.029 210 E+00
chain (66 survey feet)	meter (m)	2.011 684 E+01
statute mile (5280 survey feet)	meter (m)	1.609 347 E+03

Other Unit Conversions of Length

1 inch	=	**1.0** **E+03** mils
1 furlong	=	**660** survey feet
1 light year	=	5.878 63 E+12 miles (international)
	=	3.066 01 E-01 parsecs
1 mile (international nautical)	=	6.076 116 E+03 feet
1 mile (U.S. statute)	=	1.760 003 E+03 yards
	=	5.280 010 E+03 feet
	=	6.336 012 E+04 inches
	=	8.689 779 E-01 miles (nautical)

[238] The international nautical mile of 1.852E+03 meters was adopted in 1954 in the U.S.
[239] All conversion factors for units of land measure in this section are based on the U.S. survey foot and the following relationships:

1 rod (pole or perch) = 25 links
1 chain = 4 rods = 100 links
1 mile (U.S. statute) = 80 chains = 320 rods = 8000 links
1 acre = 160 square rods
1 square mile (U.S. statute) = 640 acres
1 section of land = 1 mile square
1 township = 6 miles square = 36 sections = 36 square miles

Area

The SI unit of area is the meter squared (m^2). Conversions for U.S. Survey units are given in a separate table, below.

Conversion to square meters from other units

To convert from	to	Multiply by	
barn	square meter (m^2)	**1.0**	**E-28**
circular mil[240]	square meter (m^2)	5.067 075 E-10	
darcy[241]	square meter (m^2)	9.869 233 E-13	
hectare	square meter (m^2)	**1.0**	**E+04**
square foot	square meter (m^2)	**9.290 304**	**E-02**
square inch	square meter (m^2)	**6.451 6**	**E-04**
square mile	square meter (m^2)	2.589 988 E+06	
square yard	square meter (m^2)	8.361 274 E-01	

Conversion to square meters from U.S. survey units

To convert from	to	Multiply by
square foot	square meter (m^2)	9.290 341 E-02
square rod	square meter (m^2)	2.529 295 E+01
square chain	square meter (m^2)	4.046 873 E+02
acre	square meter (m^2)	4.046 873 E+03
square mile	square meter (m^2)	2.589 998 E+06

Other Unit Conversions of Area

1 acre	=	**4.356**	**E+04** square survey feet
1 circular mil	=	**1.0**	**E-06** circular inches
	=	7.853 982 E-01 square mils	
1 square inch	=	1.273 240 E+06 circular mils	
	=	**6.451 6**	**E+00** square centimeters
1 square yard	=	**9.0**	**E+00** square feet

[240] A circular mil is the area of a circle 1 mil (0.001 inch) in diameter.
[241] The darcy is a unit for measuring permeability of porous solids.

Volume

The SI unit for volume is the cubic meter (m^3). Separate tables are given for U.S. units of dry volume and liquid volume, below.

Conversion to cubic meters from other units of volume

To convert from	to	Multiply by
cubic foot	cubic meter (m^3)	2.831 685 E-02
cubic inch[242]	cubic meter (m^3)	1.638 706 E-05
cubic yard	cubic meter (m^3)	7.645 549 E-01
liter[243]	cubic meter (m^3)	**1.0 E-03**
stere	cubic meter (m^3)	**1.0 E+00**

Conversion to cubic meters from other units of dry volume measure

To convert from	to	Multiply by
dry pint	cubic meter (m^3)	5.506 105 E-04
dry quart	cubic meter (m^3)	1.101 221 E-03
peck	cubic meter (m^3)	8.809 768 E-03
bushel	cubic meter (m^3)	3.523 907 E-02

Conversion to cubic meters from other units of liquid volume measure

To convert from	to	Multiply by
fluid ounce	cubic meter (m^3)	2.957 353 E-05
liquid pint	cubic meter (m^3)	4.731 765 E-04
liquid quart	cubic meter (m^3)	9.463 529 E-04
gallon	cubic meter (m^3)	3.785 412 E-03

Other Unit Conversions of Volume

1 cubic centimeter	=	6.102 374 E-02 cubic inches
	=	3.381 402 E-02 fluid ounces
1 cubic foot	=	**1.728 E+03** cubic inches
	=	2.831 685 E+01 liters
	=	7.480 520 E+00 gallons
	=	8.035 640 E-01 bushels
1 gallon	=	**2.31 E+02** cubic inches
	=	1.336 806 E-01 cubic feet
	=	3.785 412 E+00 liters
1 liter	=	**1.0 E+03** cubic centimeters
	=	3.381 402 E+01 fluid ounces
	=	1.056 688 E+00 liquid quarts
	=	2.641 721 E-01 liquid gallons
1 liquid pint	=	**0.5** liquid quart
	=	**2.887 5 E+01** cubic inches
1 liquid quart	=	**32** fluid ounces
	=	9.463 529 E-01 liters

[242] The exact conversion factor is **1.638 706 4 E-05**.

[243] In 1964 the General Conference on Weights and Measures reestablished the name liter as a special name for the cubic decimeter. Between 1901 and 1964 the liter was slightly larger (1.000 028 dm^3); in the use of high-accuracy volume data of that time interval, this fact must be kept in mind.

Mass

The SI unit of mass is the kilogram (kg). It is equal to the mass of the international prototype of the kilogram, which is a cylinder of platinum-iridium alloy kept near Paris, France. Three systems of weights exist in the U.S. and Britain: the avoirdupois, the apothecary, and the troy system. The apothecary system is used in compounding and dispensing drugs and the troy system is used for precious gems and metals. The grain is a unit common to all three systems. While the ounce and pound are used in all three systems, the definitions of these units in the avoirdupois system differ from those in the apothecary and troy systems.

Conversion to kilograms from avoirdupois units

To convert from	to	Multiply by
grain	kilogram (kg)	**6.479 891 E-05**
dram (27 11/32 grains)	kilogram (kg)	1.771 845 E-03
ounce (16 drams)	kilogram (kg)	2.834 952 E-02
pound[244] (16 oz)	kilogram (kg)	4.535 924 E-01
hundredweight , short (100 lbs)	kilogram (kg)	4.535 924 E+01
ton, net or short (2000 lbs)	kilogram (kg)	9.071 847 E+02
hundredweight, gross or long (112 lbs)	kilogram (kg)	5.080 235 E+01
ton, gross or long (2240 lbs)	kilogram (kg)	1.016 047 E+03
ton (metric)	kilogram (kg)	**1.0 E+03**
tonne	kilogram (kg)	**1.0 E+03**

Conversion to kilograms from apothecary units

To convert from	to	Multiply by
grain	kilogram (kg)	**6.479 891 E-05**
scruple (20 grains)	kilogram (kg)	1.295 978 E-03
dram (3 scruples)	kilogram (kg)	3.887 935 E-03
ounce (8 drams)	kilogram (kg)	3.110 348 E-02
pound (12 oz)	kilogram (kg)	3.732 417 E-01

Conversion to kilograms from troy units

To convert from	to	Multiply by
grain	kilogram (kg)	**6.479 891 E-05**
pennyweight (24 grains)	kilogram (kg)	1.555 174 E-03
ounce (20 pennyweights)	kilogram (kg)	3.110 348 E-02
pound (12 oz)	kilogram (kg)	3.732 417 E-01

Other Unit Conversions of Mass

1 gram	=	3.527 396 E-02 ounce (avoirdupois)
	=	1.543 236 E+01 grains

[244] The exact conversion factor is **4.535 923 7 E-01**.

Time

The SI unit of time is the second (s). The SI second is now defined as the duration of 9,192,631,770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the cesium-133 atom.

Before the present SI unit of time was adopted, other definitions were in use, including mean solar (or tropical) time and sidereal time. The mean solar day is the average elapsed time between two consecutive overhead passes of the Sun (transit of the Sun across the prime meridian). Because of the Earth's uneven rotation, and small irregularities in the direction of the Earth's axis, the mean solar day is not an invariable unit of time. A mean solar year is the time between successive passes of the Sun through the vernal equinox (marking the beginning of spring in the northern hemisphere and defined by the intersection of the ecliptic or plane of the Earth's orbit about the Sun and the celestial equator or plane through the Earth's center, perpendicular to the Earth's spin) and is 365.242 19 mean solar days long.

In sidereal (astral) time, the "fixed" (remote) stars form the frame of reference. A sidereal day is a measure of the elapsed time between two consecutive overhead passes of a position in the sky fixed by the stars. This is about four minutes less than a mean solar day, primarily because the direction between the Earth and the Sun rotates as the Earth moves about its orbit. A sidereal year represents the time required for the mean Sun to return to the same position in the sky relative to the distant stars and is 365.256 36 mean solar days long. The sidereal year is slightly longer than the mean solar year because the vernal equinox (the reference for the mean solar year) is moving with respect to the remote stars as a result of the precession of the Earth's axis of rotation.

Presently, Universal Coordinated Time (UTC) which is based on the SI second, is kept in step with mean solar time by the addition of a second (a leap second) to the UTC day whenever necessary.

Conversion to seconds from other units

To convert from	to	Multiply by
minute	seconds (s)	**6.0 E+01**
hour	seconds (s)	**3.6 E+03**
day	seconds (s)	**8.64 E+04**
year (365 days)	seconds (s)	**3.153 6 E+07**
second (sidereal)	seconds (s)	9.972 696 E-01
minute (sidereal)	seconds (s)	5.983 617 E+01
hour (sidereal)	seconds (s)	3.590 170 E+03
day (sidereal)	seconds (s)	8.616 409 E+04
year (sidereal)	seconds (s)	3.155 815 E+07
year (mean solar or tropical)	seconds (s)	3.155 693 E+07

Other Unit Conversions of Time

1 day (mean solar)	=	**24** hours (mean solar)
	=	1.002 738 E+00 days (sidereal)
	=	2.406 571 0 E+01 hours (sidereal)
1 day (sidereal)	=	**24** hours (sidereal)
	=	**1440** minutes (sidereal)
	=	1.436 068 2 E+03 minutes (mean solar)
1 week (calendar)	=	**7** days (mean solar)
	=	**168** hours (mean solar)
	=	7.019 165 4 E+00 days (sidereal)
	=	1.684 599 7 E+02 hours (sidereal)
1 year (calendar)	=	**365** days (mean solar)
	=	8760 hours (mean solar)
1 year (mean solar)	=	3.652 421 9 E+02 days (mean solar)
	=	8.765 812 6 E+03 hours (mean solar)
	=	3.662 421 9 E+02 days (sidereal)
	=	8.789 812 6 E+03 hours (sidereal)
1 year (sidereal)	=	3.652 563 6 E+02 days (mean solar)
	=	3.662 564 0 E+02 days (sidereal)

Density

The SI unit for mass per unit volume (density) is kilograms per cubic meter (kg/m^3).

Conversion to kilograms per cubic meter from other units

To convert from	to	Multiply by
grams per cubic centimeter	kilogram per cubic meter	**1.0 E+03**
pounds per cubic foot	kilogram per cubic meter	1.601 846 E+01
pounds per cubic inch	kilogram per cubic meter	2.767 990 E+04

Other Unit Conversions of Density

1 gram per cubic centimeter	=	6.242 796 E+01 pounds per cubic foot
	=	3.612 729 E-02 pounds per cubic inch
1 pound per cubic foot	=	5.787 037 E-04 pounds per cubic inch

Velocity

The SI unit of velocity is meter per second (m/s).

Conversion to meter per second from other units

To convert from	to	Multiply by
foot per hour	meter per second (m/s)	8.466 667 E-05
foot per minute	meter per second (m/s)	**5.08 E-03**
foot per second	meter per second (m/s)	**3.048 E-01**
inch per second	meter per second (m/s)	**2.54 E-02**
kilometer per hour	meter per second (m/s)	2.777 778 E-01
knot (international)	meter per second (m/s)	5.144 444 E-01
mile per hour (international)	meter per second (m/s)	**4.470 4 E-01**
mile per minute (international)	meter per second (m/s)	**2.682 24 E+01**
mile per second (international)	meter per second (m/s)	**1.609 344 E+03**

Other Unit Conversions of Velocity

1 meter per second	=	2.236 936 E+00 miles (international) per hour
1 mile (international) per hour	=	1.466 667 E+00 feet per second
	=	8.689 762 E-01 knots (international)
	=	2.777 778 E-04 miles (international) per second
	=	**1.609 344 E+00** kilometer per hour (km/h) [24f]
1 knot (international)	=	**1** mile (international nautical) per hour
	=	1.687 810 E+00 feet per second

[24f] Although speedometers may read km/h, the SI unit is m/s.

Angular Velocity

The SI unit of angular velocity is radian per second (rad/s).

Conversion to radian per second from other units

To convert from	to	Multiply by
revolution per second	radian per second (rad/s)	6.283 185 E+00
revolution per minute (RPM)	radian per second (rad/s)	1.047 198 E-01
degree per second	radian per second (rad/s)	1.745 329 E-02

Other Unit Conversions of Angular Velocity

1 RPM = **6** degrees per second
= 1.666 667 E-02 revolutions per second

Acceleration

The SI unit of acceleration is the meter per second squared (m/s^2). The SI unit of angular acceleration is radians per second squared (rad/s^2).

Conversion to meters per second squared from other units

To convert from	to	Multiply by	
foot per second squared	meter per second squared (m/s^2)	**3.048**	**E-01**
free fall, standard (g)	meter per second squared (m/s^2)	**9.806 65**	**E+00**
gal[246]	meter per second squared (m/s^2)	**1.0**	**E-02**
inch per second squared	meter per second squared (m/s^2)	**2.54**	**E-02**

Other Unit Conversions of Acceleration

1 foot per second squared = **3.048** **E+01** centimeter per second squared

[246] The unit gal, not to be confused with the gallon, refers to the local acceleration due to gravity, and is 1 centimeter per second squared.

Force

The SI Derived Unit of force is the newton (N), which is the force which, when applied to a body having a mass of one kilogram, gives it an acceleration of one meter per second squared. The principal departure of SI from the gravimetric system of metric engineering units is the use of explicitly distinct units for mass and force. In SI, the name kilogram is restricted to the unit of mass and the kilogram-force (from which the suffix force was in practice often erroneously dropped) should not be used. In its place the SI unit of force, the newton, is used.

Conversion to newton from other units

To convert from	to	Multiply by	
dyne	newton (N)	**1.0**	**E-05**
kilogram-force	newton (N)	**9.806 65**	**E+00**
kilopond (kp)	newton (N)	**9.806 65**	**E+00**
kip (1000 lbf)	newton (N)	4.448 222	E+03
ounce-force	newton (N)	2.780 139	E-01
poundal	newton (N)	1.382 550	E-01
pound-force (lbf)[247]	newton (N)	4.448 222	E+00

Other Unit Conversions of Force

1 newton = **1.0** **E+05** dyne
= 2.248 089 E-01 pound-force
1 dyne = 2.248 089 E-06 pound-force
= 7.233 012 E-05 poundal
= 1.019 716 E-03 gram-force
1 gram weight, now called gram-force = 7.093 161 E-02 poundal
= 9.806 650 E+02 dynes
= 2.204 622 E-03 pound-force
1 pound (avoirdupois) weight, now called pound-force = 3.217 404 E+01 poundal
= 4.535 924 E+02 gram-force
1 poundal = 3.108 096 E-02 pound-force
= 1.409 809 E+01 gram-force
= 1.382 550 E+04 dyne
1 lbf/lb (thrust/weight [mass] ratio) = **9.806 65** **E+00** newton per kilogram (N/kg)

[247] The exact conversion factor is **4.448 221 615 260 5 E+00**. Pound-force was formerly called *pound* (avoirdupois) weight. Gram-force was formerly called *gram* weight.

Work and Energy

The SI Derived Unit of energy is the joule (J), which is the work done when the point of application of a force of one newton is displaced a distance of one meter in the direction of the force. The units of the joule are kg $m^2 s^{-2}$.

Conversion to joules from other units

To convert from	to	Multiply by
Btu (International Table)[248]	joule (J)	1.055 056 E+03
Btu (mean)	joule (J)	1.055 87 E+03
Btu (thermochemical)	joule (J)	1.054 350 E+03
Btu (39°F)	joule (J)	1.059 67 E+03
Btu (59°F)	joule (J)	1.054 80 E+03
Btu (60°F)	joule (J)	1.054 68 E+03
calorie (International Table)	joule (J)	**4.186 8 E+00**
calorie (mean)	joule (J)	4.190 02 E+00
calorie (thermochemical)	joule (J)	**4.184 E+00**
calorie (15°C)	joule (J)	4.185 80 E+00
calorie (20°C)	joule (J)	4.181 90 E+00
calorie (kilogram, International Table)	joule (J)	**4.186 8 E+03**
calorie (kilogram, mean)	joule (J)	4.190 02 E+03
calorie (kilogram, thermochemical)	joule (J)	**4.184 E+03**
electronvolt	joule (J)	1.602 177 E-19
erg	joule (J)	**1.0 E-07**
foot poundal	joule (J)	4.214 011 E-02
foot pound-force	joule (J)	1.355 818 E+00
kilowatt hour	joule (J)	**3.6 E+06**
therm (European Community)[249]	joule (J)	**1.055 06 E+08**
therm (U.S.)[249]	joule (J)	**1.054 804 E+08**
ton (nuclear equivalent of TNT)	joule (J)	**4.184 E+09**
watt hour	joule (J)	**3.6 E+03**
watt second	joule (J)	**1.0 E+00**

Other Unit Conversions of Energy

1 joule = **1** newton meter
= **1.0 E+07** erg
= **1.0 E+07** dyne centimeter
= **1** watt second
= **1** volt coulomb
= 7.375 621 E-01 foot pound-force
= 2.386 624 E-04 calorie (kilogram, mean)
= 9.470 86 E-04 Btu (mean)
= 2.373 036 E+01 foot poundal
= 2.777 778 E-07 kilowatt-hour
= 3.725 088 E-07 horsepower (electric) hour

1 joule (International) = 1.000 165 E+00 joule[250]

[248] The Fifth International Conference on the Properties of Steam in 1956 defined the calorie (International Table) as 4.1868 J. Therefore, the exact conversion for Btu (International Table) is **1.055 055 852 62 E+03** J.

[249] The therm (European Community) is legally defined in the Council of the European Communities Directive 80 [8] EC of December 20, 1979. The therm (U.S.) is legally defined in the Federal Register, Vol. 33, No. 146, p. 10756, of July 27, 1968. Although the European therm, which is based on the International Table Btu, is frequently used by engineers in the U.S., the therm (U.S.) is the legal unit used by the U.S. gas industry.

[250] The SI Derived Unit, the joule, was formerly called the absolute joule, to distinguish it from the International Table joule.

Power

The SI unit of power is the watt (W or kg $m^2 s^{-3}$). The watt is the power which gives rise to the production of energy at the rate of one joule per second (Js^{-1}).

Conversion to watts from other units

To convert from	to	Multiply by
Btu (mean) per second	watt (W)	1.055 87 E+03
Btu (International Table) per hour	watt (W)	2.930 711 E-01
Btu (International Table) per second	watt (W)	1.055 056 E+03
Btu (thermochemical) per hour	watt (W)	2.928 751 E-01
Btu (thermochemical) per minute	watt (W)	1.757 250 E+01
Btu (thermochemical) per second	watt (W)	1.054 350 E+03
calorie (thermochemical) per minute	watt (W)	6.973 333 E-02
calorie (thermochemical) per second	watt (W)	**4.184 E+00**
erg per second	watt (W)	**1.0 E-07**
foot pound-force per hour	watt (W)	3.766 161 E-04
foot pound-force per minute	watt (W)	2.259 697 E-02
foot pound-force per second	watt (W)	1.355 818 E+00
horsepower (550 ft lbf/s)	watt (W)	7.456 999 E+02
horsepower (boiler)	watt (W)	9.809 50 E+03
horsepower (electric)	watt (W)	**7.46 E+02**
horsepower (metric)	watt (W)	7.354 988 E+02
horsepower (water)	watt (W)	7.460 43 E+02
horsepower (U.K.)	watt (W)	7.457 0 E+02
kilocalorie (thermochemical) per minute	watt (W)	6.973 333 E+01
kilocalorie (thermochemical) per second	watt (W)	**4.184 E+03**
ton of refrigeration (=12,000 Btu/hour)	watt (W)	3.516 853 E+03

Other Unit Conversions of Power

1 Btu (International Table) per minute = 2.357 141 E-02 horsepower (electric)
= 1.758 427 E+01 watt
= 7.781 694 E+02 foot pound-force per minute

1 erg per second = 5.682 518 E-09 Btu (mean) per minute
= 1.431 97 E-09 kilogram calorie (mean) per minute
= 7.375 621 E-08 foot pound-force per second

1 horsepower (U.K.) = 7.062 43 E-01 Btu (mean) per second
= **550** foot pound-force per second

1 quad = **1.0 E+15** Btu

Torque

The SI unit for torque (or bending moment) is newton meter (N·m).[251]

Conversion to newton meter from other units

To convert from	to	Multiply by	
dyne centimeter	newton meter (N·m)	**1.0**	**E-07**
kilogram-force[252] meter	newton meter (N·m)	**9.806 65**	**E+00**
ounce-force inch	newton meter (N·m)	7.061 552	E-03
pound-force inch	newton meter (N·m)	1.129 848	E-01
pound-force foot	newton meter (N·m)	1.355 818	E+00

Other Unit Conversions of Torque

1 dyne centimeter	=	1.019 716 E-08 kilogram-force meter
	=	7.375 621 E-08 pound-force foot
	=	8.850 748 E-07 pound-force inch
	=	2.373 036 E-06 poundal foot
1 kilogram-force meter	=	**9.806 65 E+07** dyne centimeter
1 pound-force foot	=	1.355 818 E+07 dyne centimeter
1 poundal foot	=	4.214 012 E+05 dyne centimeter
1 pound-force inch	=	1.129 848 E+06 dyne centimeter

[251] The vector product of force and moment arm is widely designated by the unit newton meter. This unit for bending moment or torque results in confusion with the unit for energy, which is also newton meter. If torque is expressed as newton meter per radian, the relationship to energy is clarified, since the product of torque and angular rotation is energy: (Nm/rad) rad = Nm. If the vectors were shown, the distinction between energy and torque would be obvious since the orientation of force and length is different in the two cases.

[252] See discussion under FORCE section. Considerable confusion exists in the use of the term "weight" as a quantity to mean either "force" or "mass". In commercial and everyday use, the term weight nearly always means mass; thus when one speaks of a person's weight, the quantity referred to is mass. This non-technical use of the term weight will probably persist. In science and technology, the term "weight of a body" has usually meant the force that, if applied to the body, would give it an acceleration equal to the local acceleration of free fall ("local" usually means a location on the surface of the earth).

When non-SI units are used, a distinction should be made between force and mass. For example, pound-force to denote force in gravimetric engineering units and pound for mass.

Pressure

The SI unit of pressure or stress is the pascal, Pa (newton per square meter, or N m^{-2}).

Conversion to pascal from other units

To convert from	to	Multiply by	
atmosphere, standard	pascal (Pa)	**1.013 25**	**E+05**
atmosphere, technical (1 kilogram-force/cm^2)	pascal (Pa)	**9.806 65**	**E+04**
bar	pascal (Pa)	**1.0**	**E+05**
centimeter of mercury (0 °C)	pascal (Pa)	1.333 22	E+03
centimeter of water (4 °C)	pascal (Pa)	9.806 38	E+01
dyne per cm^2	pascal (Pa)	**1.0**	**E-01**
foot of water (39.2 °F)	pascal (Pa)	2.988 98	E+03
gram-force per cm^2	pascal (Pa)	**9.806 65**	**E+01**
inch of mercury (32 °F)	pascal (Pa)	3.386 38	E+03
inch of mercury (60 °F)	pascal (Pa)	3.376 85	E+03
inch of water (39.2 °F)	pascal (Pa)	2.490 82	E+02
inch of water (60 °F)	pascal (Pa)	2.488 4	E+02
kilogram-force per cm^2	pascal (Pa)	**9.806 65**	**E+04**
kilogram-force per m^2	pascal (Pa)	**9.806 65**	**E+00**
kilogram-force per mm^2	pascal (Pa)	**9.806 65**	**E+06**
kip (1000 pound-force) per in^2 (ksi)	pascal (Pa)	6.894 757	E+06
millibar	pascal (Pa)	**1.0**	**E+02**
millimeter of mercury (0 °C)	pascal (Pa)	1.333 22	E+02
poundal per ft^2	pascal (Pa)	1.488 164	E+00
pound-force per ft^2	pascal (Pa)	4.788 026	E+01
pound-force per in^2 (psi)	pascal (Pa)	6.894 757	E+03
torr (mm Hg, 0 °C)	pascal (Pa)	1.333 22	E+02

Other Unit Conversions of Pressure or Stress

1 standard atmosphere	=	1.013 250 E+00 bars
	=	1.469 595 E+01 pound-force per in^2 (psi)
	=	1.013 250 E+06 dynes per cm^2
	=	**7.60 E+02** torr (mm Hg, 0 °C)
	=	2.992 13 E+01 inches Hg, 32 °F
	=	3.389 95 E+01 feet of water (39.2 °F)
1 dyne per cm^2	=	1.450 377 E-05 pound-force per in^2 (psi)
1 bar	=	**1.0 E+06** dynes per cm^2
	=	9.869 233 E-01 standard atmosphere
1 pound-force per in^2	=	6.894 757 E+04 dynes per cm^2

Temperature

The SI unit of temperature is the kelvin (K). The kelvin, a unit of thermodynamic temperature, is the fraction 1/273.16 of the thermodynamic temperature of the triple point of water.[253]

Wide use is also made of the degree Celsius (°C), which is the SI unit for expressing Celsius temperature and temperature intervals. The Celsius scale (formerly called centigrade) is related directly to thermodynamic temperature (kelvin) as follows:

The temperature interval one degree Celsius equals one kelvin exactly.

Celsius temperature (t) is related to thermodynamic temperature (T) by the equation:

$$t = T - T_0$$

where T_0 = **273.15** K by definition.

To convert from degrees Fahrenheit to degrees Celsius:

$$t_C = (t_F - 32)/1.8$$

Conversion to kelvin from other temperature units

To convert from	to	Use
degree Celsius	kelvin (K)	$T_K = t_C + 273.15$
degree Fahrenheit	kelvin (K)	$T_K = (t_F + 459.67)/1.8$
degree Rankine	kelvin (K)	$T_K = T_{°R}/1.8$

Temperature Differences

Δ1° Celsius (C)	=	Δ**1.8**° Fahrenheit (F)
	=	Δ**1.8**° Rankine (R)
	=	Δ**1** kelvin (K)

Other Unit Conversions of Temperature

0°C	=	**273.15** K
	=	**32**°F
	=	**491.67**°R
100°C	=	**373.15** K
	=	**212**°F
	=	**671.67**°R
0K	=	**-273.15**°C
	=	**-459.67**°F

[253] The triple point of water is 273.16K (0.01°C), which is 0.01 K above the freezing point of water at a pressure of 101.325 kPa.

Electrical Units

The SI Base Unit of electric current is the ampere (A). The ampere is that constant current which, if maintained in two straight parallel conductors of infinite length, of negligible circular cross-section, and placed one meter apart in a vacuum, would produce between these conductors a force equal to 2×10^{-7} newton per meter of length.

The SI Derived Unit for quantity of electricity is the coulomb (C), which is the quantity of electricity transported in one second by a current of one ampere. Other SI Derived Units for electrical quantities include that for electric capacitance: the farad (F), which is the capacitance of a capacitor between the plates of which there appears a difference of potential of one volt when it is charged by a quantity of electricity equal to one coulomb. Electrical conductance: the siemens (S), which is the electric conductance of a conductor in which a current of one ampere is produced by an electric potential difference of one volt. Electric inductance: the henry (H), which is the inductance of a closed circuit in which an electromotive force of one volt is produced when the electric current in the circuit varies uniformly at the rate of one ampere per second. Electromotive force or electric potential difference: the volt (V), which is the difference of electric potential between two points of a conductor carrying a constant current of one ampere, when the power dissipated between these points is equal to one watt. Electric resistance: the ohm (Ω), which is the electric resistance between two points of a conductor when a constant difference of potential of one volt, applied between these two points, produces in this conductor a current of one ampere, this conductor not being the source of any electromotive force.

Conversion to SI electrical units from other units

To convert from	to	Multiply by	
abampere	ampere (A)	**1.0**	**E+01**
EMU of current	ampere (A)	**1.0**	**E+01**
ESU of current	ampere (A)	3.335 641	E-10
gilbert	ampere (A)	7.957 747	E-01
statampere	ampere (A)	3.335 641	E-10
oersted	ampere per meter (A/m)	7.957 747	E+01
abcoulomb	coulomb (C)	**1.0**	**E+01**
ampere hour	coulomb (C)	**3.6**	**E+03**
faraday (based on C-12)	coulomb (C)	9.648 531	E+04
faraday (chemical)	coulomb (C)	9.649 57	E+04
faraday (physical)	coulomb (C)	9.652 19	E+04
statcoulomb	coulomb (C)	3.335 641	E-10
abfarad	farad (F)	**1.0**	**E+09**
EMU of capacitance[254]	farad (F)	**1.0**	**E+09**
ESU of capacitance	farad (F)	1.112 650	E-12
statfarad	farad (F)	1.112 650	E-12
abhenry	henry (H)	**1.0**	**E-09**
EMU of inductance	henry (H)	**1.0**	**E-09**
ESU of inductance	henry (H)	8.987 552	E+11
stathenry	henry (H)	8.987 552	E+11
abohm	ohm (Ω)	**1.0**	**E-09**
EMU of resistance	ohm (Ω)	**1.0**	**E-09**
ESU of resistance	ohm (Ω)	8.987 552	E+11
statohm	ohm (Ω)	8.987 552	E+11
abmho	siemens (S)	**1.0**	**E+09**
mho	siemens (S)	**1.0**	**E+00**
statmho	siemens (S)	1.112 650	E-12
abvolt	volt (V)	**1.0**	**E-08**
EMU of electric potential	volt (V)	**1.0**	**E-08**
ESU of electric potential	volt (V)	2.997 925	E+02
statvolt	volt (V)	2.997 925	E+02

[254] EMU means electromagnetic cgs unit. ESU means electrostatic cgs unit.

Magnetic Units

The SI Derived Unit of magnetic flux is the weber (Wb): the magnetic flux which, linking a circuit of one turn, produces in it an electromotive force of one volt as it is reduced to zero at a uniform rate in one second. The SI Derived Unit for magnetic flux density is the tesla (T), which is the magnetic flux density given by a magnetic flux of one weber per square meter.

Conversion to SI magnetic units from other units

To convert from	to	Multiply by	
maxwell	weber (Wb)	**1.0**	**E-08**
unit pole	weber (Wb)	1.256 637	E-07
gamma	tesla (T)	**1.0**	**E-09**
gauss	tesla (T)	**1.0**	**E-04**

Radiation Units

The SI Derived Unit of absorbed dose is the gray[255] (Gy) which is the absorbed dose when the energy per unit mass imparted to matter by ionizing radiation is one joule per kilogram. The SI Derived Unit of dose equivalent is the sievert (Sv), which is the dose equivalent when the absorbed dose of ionizing radiation multiplied by the dimensionless factors Q (Quality factor) and N (product of any other multiplying factors) stipulated by the International Commission on Radiological Protection (ICRP) is one joule per kilogram. The SI Derived Unit of activity is the becquerel (Bq), which is the activity of a radionuclide decaying at the rate of one spontaneous nuclear transition per second.

Conversion to SI radiation units from other units

To convert from	to	Multiply by	
curie	becquerel (Bq)	**3.7**	**E+10**
rad	gray (Gy)	**1.0**	**E-02**
rem	sievert (Sv)	**1.0**	**E-02**
roentgen	coulomb per kilogram (C/kg)	**2.58**	**E-04**

[255] The gray is also used for the ionizing radiation quantities specific energy imparted, kerma, and absorbed dose index, which have the SI unit joule per kilogram.

Other SI Units

The SI Base Unit for amount of substance is the mole (mol), the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilograms of carbon-12. When the mole is used, the elementary entities must be specified and may be atoms, molecules, ions, electrons, other particles, or specified group of particles.

The SI Derived Unit of frequency is the hertz (Hz), which is the frequency of a periodic phenomenon of which the period is one second.

The SI Derived Unit of luminous flux is the lumen (lm), which is the luminous flux emitted in a solid angle of one steradian by a point source having a uniform intensity of one candela.

The SI Derived Unit of illuminance is the lux (lx), which is the illuminance produced by a luminous flux of one lumen uniformly distributed over a surface of one square meter.

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Alphabetical List of the Elements

Name	Sym	Z	Nat. Occ. Isotopes*	Mass (u)	Name	Sym	Z	Nat. Occ. Isotopes*	Mass (u)
actinium	Ac	89		.....	meitnerium	Mt	109		.....
aluminum	Al	13	1	26.981 538(2)	mendelevium	Md	101		.....
americium	Am	95		.....	mercury	Hg	80	7	200.59(2)
antimony	Sb	51	2	121.760(1)	molybdenum	Mo	42	7	95.94(2)
argon	Ar	18	3	39.948(1)	neodymium	Nd	60	7	144.24(3)
arsenic	As	33	1	74.921 60(2)	neon	Ne	10	3	20.1797(6)
astatine	At	85		.....	neptunium	Np	93		.....
barium	Ba	56	7	137.327(7)	nickel	Ni	28	5	58.6934(2)
berkelium	Bk	97		.....	niobium	Nb	41	1	92.906 38(2)
beryllium	Be	4	1	9.012 182(3)	nitrogen	N	7	2	14.0067(2)
bismuth	Bi	83	1	208.980 38(2)	nobelium	No	102		.....
bohrium	Bh	107		...	osmium	Os	76	7	190.23(3)
boron	B	5	2	10.811(7)	oxygen	O	8	3	15.9994(3)
bromine	Br	35	2	79.904(1)	palladium	Pd	46	6	106.42(1)
cadmium	Cd	48	8	112.411(8)	phosphorus	P	15	1	30.973 761(2)
calcium	Ca	20	6	40.078(4)	platinum	Pt	78	6	195.078(2)
californium	Cf	98		.....	plutonium	Pu	94		.....
carbon	C	6	2	12.0107(8)	polonium	Po	84		.....
cerium	Ce	58	4	140.116(1)	potassium	K	19	3	39.0983(1)
cesium	Cs	55	1	132.905 45(2)	praseodymium	Pr	59	1	140.907 65(2)
chlorine	Cl	17	2	35.453(2)	promethium	Pm	61		.....
chromium	Cr	24	4	51.9961(6)	protactinium	Pa	91	1	231.035 88(2)
cobalt	Co	27	1	58.933 200(9)	radium	Ra	88		.....
copper	Cu	29	2	63.546(3)	radon	Rn	86		.....
curium	Cm	96		.....	rhenium	Re	75	2	186.207(1)
darmstadtium	Ds	110		.....	rhodium	Rh	45	1	102.905 50(2)
dubnium	Db	105		.....	rubidium	Rb	37	2	85.4678(3)
dysprosium	Dy	66	7	162.500(1)	ruthenium	Ru	44	7	101.07(2)
einsteinium	Es	99		.....	rutherfordium	Rf	104		.....
erbium	Er	68	6	167.259(3)	samarium	Sm	62	7	150.36(3)
europium	Eu	63	2	151.964(1)	scandium	Sc	21	1	44.955 910(8)
fermium	Fm	100		.....	seaborgium	Sg	106		.....
fluorine	F	9	1	18.998 4032(5)	selenium	Se	34	6	78.96(3)
francium	Fr	87		.....	silicon	Si	14	3	28.0855(3)
gadolinium	Gd	64	7	157.25(3)	silver	Ag	47	2	107.8682(2)
gallium	Ga	31	2	69.723(1)	sodium	Na	11	1	22.989 770(2)
germanium	Ge	32	5	72.64(1)	strontium	Sr	38	4	87.62(1)
gold	Au	79	1	196.966 55(2)	sulfur	S	16	4	32.065(5)
hafnium	Hf	72	6	178.49(2)	tantalum	Ta	73	2	180.9479(1)
hassium	Hs	108		.....	technetium	Tc	43		.....
helium	He	2	2	4.002 602(2)	tellurium	Te	52	8	127.60(3)
holmium	Ho	67	1	164.930 32(2)	terbium	Tb	65	1	158.925 34(2)
hydrogen	H	1	2	1.007 94(7)	thallium	Tl	81	2	204.3833(2)
indium	In	49	2	114.818(3)	thorium	Th	90	1	232.0381(1)
iodine	I	53	1	126.904 47(3)	thulium	Tm	69	1	168.934 21(2)
iridium	Ir	77	2	192.217(3)	tin	Sn	50	10	118.710(7)
iron	Fe	26	4	55.845(2)	titanium	Ti	22	5	47.867(1)
krypton	Kr	36	6	83.798(2)	tungsten	W	74	5	183.84(1)
lanthanum	La	57	2	138.9055(2)	uranium	U	92	3	238.028 91(3)
lawrencium	Lr	103		.....	vanadium	V	23	2	50.9415(1)
lead	Pb	82	4	207.2(1)	xenon	Xe	54	9	131.293(6)
lithium	Li	3	2	6.941(2)	ytterbium	Yb	70	7	173.04(3)
lutetium	Lu	71	2	174.967(1)	yttrium	Y	39	1	88.905 85(2)
magnesium	Mg	12	3	24.3050(6)	zinc	Zn	30	5	65.409(4)
manganese	Mn	25	1	54.938 049(9)	zirconium	Zr	40	5	91.224(2)

* The number of naturally-occurring isotopes.

Periodic Table of the Elements

Knolls Atomic Power Laboratory

Symbol ↓

Atomic Number → 92 **U** ← Oxidation States—most common in red

← Electronegativity (Pauling scale)

238.0291 ← Atomic Mass

Uranium ← Name

Melting Point (°C) → 1135 ← Atomic Radius (pm)

Boiling Point (°C) → 4131 ← K or L α X-ray energy (keV)

Specific Gravity–(density g/l for gases at 0°C) → 18.95 ← K or L β X-ray energy (keV)

Thermal Conductivity (W/(m K)) → 27.6 ← crustal abundance (ppm by mass)

Note: Color of Symbol and Name indicates gas, liquid, **solid** (at 20°C) or artificially prepared;
"d" on C (Z=6) represents "diamond" and "w" on P (Z=15) represents "white phosphorus"

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.00794 Hydrogen -259.34 / -252.87 / 0.0899 / 0.1815 / 1400 / 76																	2 He 4.002602 Helium -272.2 / -268.93 / 0.1785 / 0.152 / 0.008 / 128
3 Li 6.941 Lithium 180.5 / 1342 / 0.534 / 84.7 / 20 / 1.0 / 0.98 / 152	4 Be 9.012182 Beryllium 1287 / 2471 / 1.848 / 200 / 2.8 / 1.57 / 113.3											5 B 10.811 Boron 2075 / 4000 / 2.36 / 27.0 / 10 / 3 / 2.04 / 83	6 C 12.0107 Carbon 3550 d / 3830 sub / 1.8-3.5 / 5-2300 / 200 / 2,4 / 2.55 / 77	7 N 14.0067 Nitrogen -210.0 / -195.79 / 0.0008 / 0.02588 / 19 / -3, 5 / 3.04 / 71	8 O 15.9994 Oxygen -218.79 / -183.95 / 1.429 / 0.0274 / 461000 / -2 / 3.44 / 66	9 F 18.9984032 Fluorine -219.67 / -188.12 / 1.696 / 0.0279 / 585 / -1 / 3.98 / 70.9	10 Ne 20.1797 Neon -248.59 / -246.08 / 0.8999 / 0.0493 / 0.005 / - / 0.849
11 Na 22.989770 Sodium 97.80 / 883 / 0.971 / 14.1 / 23600 / 1 / 0.93 / 144.4	12 Mg 24.3050 Magnesium 650 / 1090 / 1.738 / 156 / 23300 / 2 / 1.31 / 160											13 Al 26.981538 Aluminum 660.32 / 2519 / 2.6989 / 237 / 82300 / 3 / 1.61 / 143.1	14 Si 28.0855 Silicon 1414 / 3265 / 2.33 / 148 / 282000 / 4 / 1.90 / 117	15 P 30.973761 Phosphorus 44.15 w / 280.5 w / 1.82 w / 0.235 w / 1050 / 3, 5 / 2.19 / 83	16 S 32.065 Sulfur 119.6 / 444.60 / 2.07 / 0.24 / 350 / 6, 4, 2, -2 / 2.58 / 104	17 Cl 35.453 Chlorine -101.5 / -34.4 / 3.214 / 0.0089 / 145 / 7, 6, 1, -1 / 3.16 / 99	18 Ar 39.948 Argon -189.35 / -185.85 / 1.784 / 0.0177 / 3.5 / - / - / 174
19 K 39.0983 Potassium 63.5 / 759 / 0.862 / 102.4 / 20900 / 227 / 1 / 0.82	20 Ca 40.078 Calcium 842 / 1484 / 1.55 / 200 / 41500 / 197.3 / 2 / 1.00	21 Sc 44.955910 Scandium 1541 / 2836 / 2.989 / 15.8 / 22 / 165.8 / 3 / 1.36	22 Ti 47.867 Titanium 1668 / 3287 / 4.54 / 21.9 / 9650 / 144.8 / 4,3,2 / 1.54	23 V 50.9415 Vanadium 1910 / 3407 / 6.11 / 30.7 / 120 / 132.1 / 5,4,3,2 / 1.63	24 Cr 51.9961 Chromium 1907 / 2671 / 7.19 / 93.7 / 102 / 124.9 / 6,3,2 / 1.66	25 Mn 54.938049 Manganese 1246 / 2061 / 7.3 / 7.82 / 950 / 124 / 7,4,3,2 / 1.55	26 Fe 55.845 Iron 1538 / 2861 / 7.874 / 80.2 / 56300 / 124.1 / 3,2 / 1.83	27 Co 58.933200 Cobalt 1495 / 2927 / 8.9 / 100 / 25 / 124.3 / 3,2 / 1.88	28 Ni 58.6934 Nickel 1455 / 2913 / 8.902 / 90.7 / 84 / 124.6 / 3,2 / 1.91	29 Cu 63.546 Copper 1084.8 / 2562 / 8.96 / 401 / 60 / 127.8 / 2,1 / 1.90	30 Zn 65.409 Zinc 419.5 / 907 / 7.133 / 116 / 70 / 133.2 / 2 / 1.65	31 Ga 69.723 Gallium 29.76 / 2204 / 5.904 / 40.6 / 19 / 122.1 / 3 / 1.81	32 Ge 72.64 Germanium 938.25 / 2833 / 5.323 / 59.9 / 1.5 / 122.5 / 4,2 / 2.01	33 As 74.92160 Arsenic ~817 / 603 sub / 5.7 / 50.0 / 1.8 / 121 / 5,3,-3 / 2.18	34 Se 78.96 Selenium 220.5 / 685 / 4.7 / 0.9-4.5 / 0.05 / 117 / 6,4,-2 / 2.55	35 Br 79.904 Bromine -7.2 / 58.8 / 3.12 / 0.122 / 2.4 / 114 / 7,5,1,-1 / 2.96	36 Kr 83.798 Krypton -157.38 / -153.22 / 3.733 / 0.00949 / trace / 189 / - / -
37 Rb 85.4678 Rubidium 39.3 / 688 / 1.532 / 58.2 / 90 / 247.5 / 1 / 0.82	38 Sr 87.62 Strontium 777 / 1382 / 2.54 / 35.3 / 370 / 215.1 / 2 / 0.95	39 Y 88.90585 Yttrium 1522 / 3345 / 4.469 / 17.2 / 33 / 181 / 3 / 1.22	40 Zr 91.224 Zirconium 1855 / 4409 / 6.506 / 22.7 / 165 / 160 / 4 / 1.33	41 Nb 92.90638 Niobium 2477 / 4744 / 8.57 / 53.7 / 20 / 142.9 / 5,3 / 1.6	42 Mo 95.94 Molybdenum 2623 / 4639 / 10.22 / 138 / 1.2 / 136.2 / 6 / 2.16	43 Tc Technetium 2157 / 4265 / 11.50 / 50.6 / 0 / 135.6 / 7,6,4 / 2.10	44 Ru 101.07 Ruthenium 2334 / 4150 / 12.41 / 117 / 0.001 / 134 / 3 / 2.2	45 Rh 102.90550 Rhodium 1964 / 3695 / 12.41 / 150 / 0.001 / 134.5 / 3 / 2.28	46 Pd 106.42 Palladium 1554.9 / 2963 / 12.02 / 71.8 / 0.015 / 137.6 / 2,4 / 2.20	47 Ag 107.8682 Silver 961.78 / 2162 / 10.50 / 42.8 / 0.075 / 144 / 1 / 1.93	48 Cd 112.411 Cadmium 321.07 / 767 / 8.65 / 96.8 / 0.15 / 148.9 / 2 / 1.69	49 In 114.818 Indium 156.6 / 2072 / 7.31 / 81.8 / 0.25 / 162.6 / 3 / 1.78	50 Sn 118.710 Tin 231.93 / 2602 / 5.77-7.3 / 66.5 / 2.3 / 140.5 / 4,2 / 1.96	51 Sb 121.760 Antimony 630.63 / 1587 / 6.691 / 24.3 / 0.2 / 142 / 5,3,-3 / 2.05	52 Te 127.60 Tellurium 449.5 / 988 / 6.24 / 2.3 / 0.001 / 143.2 / 6,4,-2 / 2.1	53 I 126.90447 Iodine 113.7 / 184.4 / 4.93 / 0.449 / 0.45 / 133.5 / 7,5,1,-1 / 2.66	54 Xe 131.293 Xenon -111.79 / -108.12 / 5.897 / 0.00589 / trace / 218 / - / 2.60
55 Cs 132.90545 Cesium 28.5 / 671 / 1.873 / 35.9 / 3 / 265.4 / 1 / 0.79	56 Ba 137.327 Barium 727 / 1897 / 3.5 / 18.4 / 425 / 217.3 / 2 / 0.89	57 La* 138.9055 Lanthanum 918 / 3464 / 6.145 / 13.5 / 39 / 187.7 / 3 / 1.10	72 Hf 178.49 Hafnium 2233 / 4603 / 13.31 / 23.0 / 3 / 156.4 / 4 / 1.3	73 Ta 180.9479 Tantalum 3017 / 5458 / 16.654 / 57.5 / 2 / 143 / 5 / 1.5	74 W 183.84 Tungsten 3422 / 5555 / 19.3 / 174 / 1.25 / 137.0 / 6 / 1.7	75 Re 186.207 Rhenium 3186 / 5596 / 21.02 / 47.9 / 0.0007 / 137.0 / 7,6,4 / 1.9	76 Os 190.23 Osmium 3033 / 5012 / 22.57 / 87.8 / 0.0015 / 135 / 3,4 / 2.2	77 Ir 192.217 Iridium 2446 / 4428 / 22.42 / 147 / 0.001 / 135.7 / 3,4 / 2.2	78 Pt 195.078 Platinum 1768.4 / 3825 / 21.45 / 71.8 / 0.005 / 138 / 2,4 / 2.2	79 Au 196.96655 Gold 1064.18 / 2856 / -19.3 / 317 / 0.004 / 144 / 3,1 / 2.4	80 Hg 200.59 Mercury -38.83 / 356.73 / 13.546 / 8.34 / 0.085 / 160 / 2,1 / 1.9	81 Tl 204.3833 Thallium 304 / 1473 / 11.85 / 46.1 / 0.85 / 170.4 / 3,1 / 1.8	82 Pb 207.2 Lead 327.46 / 1749 / 11.35 / 35.3 / 14 / 175.0 / 4,2 / 1.8	83 Bi 208.98038 Bismuth 271.4 / 1564 / 9.747 / 7.87 / 0.0085 / 155 / 5,3 / 1.9	84 Po Polonium 254 / 962 / 9.32 / 20 / - / 167 / 6,4,2 / 2.0	85 At Astatine 302 / ~340 / - / 1.7 / - / - / 7,5,3,1,-1 / 2.2	86 Rn Radon -71 / -61.7 / 9.73 / 0.00364 / trace / - / - / -
87 Fr Francium 27 / 677 / - / -15 / - / 270 / 1 / 0.7	88 Ra Radium 700 / 1140 / 5 / 18.6 / - / 223 / 2 / 0.9	89 Ac* Actinium 1051 / 3198 / 10.07 / 12 / - / 187.8 / 3 / 1.1	104 Rf Rutherfordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110 Ds Darmstadtium	111	112		114		116		

*Lanthanides	58 Ce 140.116 Cerium 798 / 3443 / 6.770 / 11.3 / 66.5 / 182.5 / 4,3 / 1.12	59 Pr 140.90765 Praseodymium 931 / 3520 / 6.773 / 12.5 / 9.2 / 182.8 / 3 / 1.13	60 Nd 144.24 Neodymium 1021 / 3074 / 7.008 / 16.5 / 41.5 / 182.1 / 3 / 1.14	61 Pm Promethium 1042 / ~3000 / 7.264 / ~17.9 / - / 181.0 / 3,2 / -	62 Sm 150.36 Samarium 1074 / 1794 / 7.520 / 13.3 / 7.05 / 180.2 / 3,2 / 1.17	63 Eu 151.964 Europium 822 / 1529 / 5.244 / 13.9 / 2 / 204.2 / 3,2 / -	64 Gd 157.25 Gadolinium 1313 / 3273 / 7.901 / 10.6 / 6.2 / 180.2 / 3 / 1.20	65 Tb 158.92534 Terbium 1356 / 3230 / 8.230 / 11.1 / 1.2 / 178.2 / 3 / -	66 Dy 162.500 Dysprosium 1412 / 2567 / 8.551 / 10.7 / 5.2 / 177.3 / 3 / 1.22	67 Ho 164.93032 Holmium 1474 / 2700 / 8.795 / 16.2 / 1.3 / 176.6 / 3 / 1.23	68 Er 167.259 Erbium 1529 / 2868 / 9.066 / 14.3 / 3.5 / 175.7 / 3 / 1.24	69 Tm 168.93421 Thulium 1545 / 1950 / 9.321 / 16.8 / 0.52 / 174.6 / 3,2 / 1.25	70 Yb 173.04 Ytterbium 819 / 1196 / 6.93 / 35 / 3.2 / 194 / 3,2 / -	71 Lu 174.967 Lutetium 1663 / 3402 / 9.841 / 16.4 / 0.6 / 173.4 / 3 / 1.9
*Actinides	90 Th 232.0381 Thorium 1750 / 4788 / 11.72 / 54.0 / 9.6 / 179.8 / 4 / 1.3	91 Pa 231.03588 Protactinium 1572 / ~4000 / 15.37 / ~47 / trace / 163.5 / 5,4 / 1.5	92 U 238.02891 Uranium 1135 / 4131 / 18.95 / 27.6 / 2.7 / 138.5 / 6,5,4,3 / 1.7	93 Np Neptunium 644 / 3902 / 20.25 / 0.3 / - / 131 / 6,5,4,3 / 1.3	94 Pu Plutonium 640 / 3228 / 19.84 / 6.74 / - / 151 / 6,5,4,3 / 1.3	95 Am Americium 1176 / 2011 / 13.67 / ~10 / - / 184 / 6,5,4,3 / 1.3	96 Cm Curium 1345 / 3100 / 13.5 / ~10 / - / 170 / 3	97 Bk Berkelium 1050 / - / ~14 / ~10 / - / - / 4,3	98 Cf Californium 900 / - / - / ~10 / - / ~186 / 3	99 Es Einsteinium 860 / - / - / ~10 / - / ~186 / 3,2	100 Fm Fermium 1527 / - / - / ~10 / - / - / 3,2	101 Md Mendelevium 827 / - / - / ~10 / - / - / 3,2	102 No Nobelium 827 / - / - / ~10 / - / - / 3,2	103 Lr Lawrencium 1627 / - / - / ~10 / - / - / 3

Fundamental General Physics Constants

This set of fundamental constants is excerpted from a self-consistent set prepared by P.J. Mohr and B.N. Taylor, published in Journal of Physical and Chemical Reference Data **59** (1999), "CODATA Recommended Values of the Fundamental Physical Constants: 1998", and in Rev. Mod. Phys. **72** (2000). The values can also be found on the Internet at http://physics.nist.gov/cuu/Constants/index.html.

The values given in parentheses after the stated value of the constants are the one sigma uncertainties in the last digit(s) given. The first value of each constant given is in SI units. Following values are in non-SI units. Where no value is given in parentheses for the error, it may be assumed that the value is exact, except where trailing "..." are indicated.

Acceleration due to gravity (standard)
g = 9.806 65 E+00 meter per second squared
= 3.217 405... E+01 feet per second squared

Alpha particle rest mass
m_α = 6.644 655 98(58) E-27 kilogram
= 3.727 379 04(15) E+03 MeV
= 4.001 506 1747(10) E+00 u

Atomic mass unit (AMU)= $(10^{-3}$ kg mol$^{-1})/N_A$
u = 1.660 538 73(13) E-27 kilogram
= 9.314 940 13 (37) E+02 MeV

Avogadro constant
N_A = 6.022 141 99(47) E+23 "entities" (molecules, atoms, etc.) per mole

Bohr magneton= $e\hbar/2m_e$
μ_B = 9.274 008 99(37) E-26 joules per tesla

Boltzmann constant = R/N_A
k = 1.380 6503(24) E-23 joules per kelvin
= 8.617 342(15) E-05 electron volt per kelvin

Characteristic impedance of vacuum = $(\mu_0/\varepsilon_0)^{1/2}$
Z_0 = 3.767 303 134 61...E+02 ohms

Classical electron radius = $\alpha^2\alpha_0$
r_e = 2.817 940 285(31) E-15 meter

Deuteron rest mass
m_d = 3.343 583 09(26) E-27 kilogram
= 1.875 612 762(75) E+03 MeV
= 2.013 553 212 71(35) E+00 u

Electric constant = $1/\mu_0c^2$
ε_0 = 8.854 187 817...E-12 coulomb squared per joule per meter
= 8.854 187 817...E-12 farads per meter

Electron charge to mass ratio
$-e/m_e$ = 1.758 820 174(71) E+11 coulombs per kilogram

Electron rest mass
m_e = 9.109 381 88(72) E-31 kilogram
= 5.109 989 02(21) E-01 MeV
= 5.485 799 110(12) E-04 AMU

Elementary charge
e = 1.602 176 462(63) E-19 coulombs

Faraday constant
F = 9.648 534 15(39) E+04 coulombs per mole

Fine structure constant (dimensionless)
α = 7.297 352 533(27) E-03

First Bohr radius= $\alpha/4\pi R_\infty$
α_0 = 5.291 772 083(19) E-11 meter

Gravitational constant
G = 6.673(10) E-11 newton meter squared per kilogram squared
= 6.673(10) E-08 centimeters cubed per gram per second squared, or dyne centimeter squared per gram squared

Molar gas constant
R = 8.314 472(15) E+00 joules per mole per kelvin

Magnetic constant
μ_0 = (4π) E-07 henry per meter
= (4π) E-07 newton per ampere squared
1.256 637 061 4... E-06 henry per meter
= 1.674 927 16(13) E-27 kilogram
= 9.395 653 30(38) E+02 MeV
= 1.008 664 915 78(55) E+00 u

Neutron rest mass
m_n = 1.674 927 16(13) E-27 kilograms
= 9.395 653 30(38) MeV
= 1.008 664 915 78(55) u

Nuclear magneton= $e\hbar/2m_p$
μ_N = 5.050 783 17(20) E-27 joules per tesla

Planck constant
h = 6.626 068 76(52)E-34 joule seconds or joule per Hertz
= 4.135 667 27(16) E-21 MeV seconds

Proton mass to electron mass ratio
m_p/m_e = 1.836 152 6675(39) E+03

Proton rest mass
m_p = 1.672 621 58(13) E-27 kilogram
= 9.382 719 98(38) E+02 MeV
= 1.007 276 466 88(13) E+00 u

Rydberg constant
R_∞ = 1.097 373 156 8549(83) E+07 per meter

Speed of light
c = 2.997 924 58 E+08 meters per second
= 2.997 924 58 E+10 centimeters per second

Stefan-Boltzmann constant
σ = 5.670 400(40) E-08 watts per meter squared per kelvin to fourth power

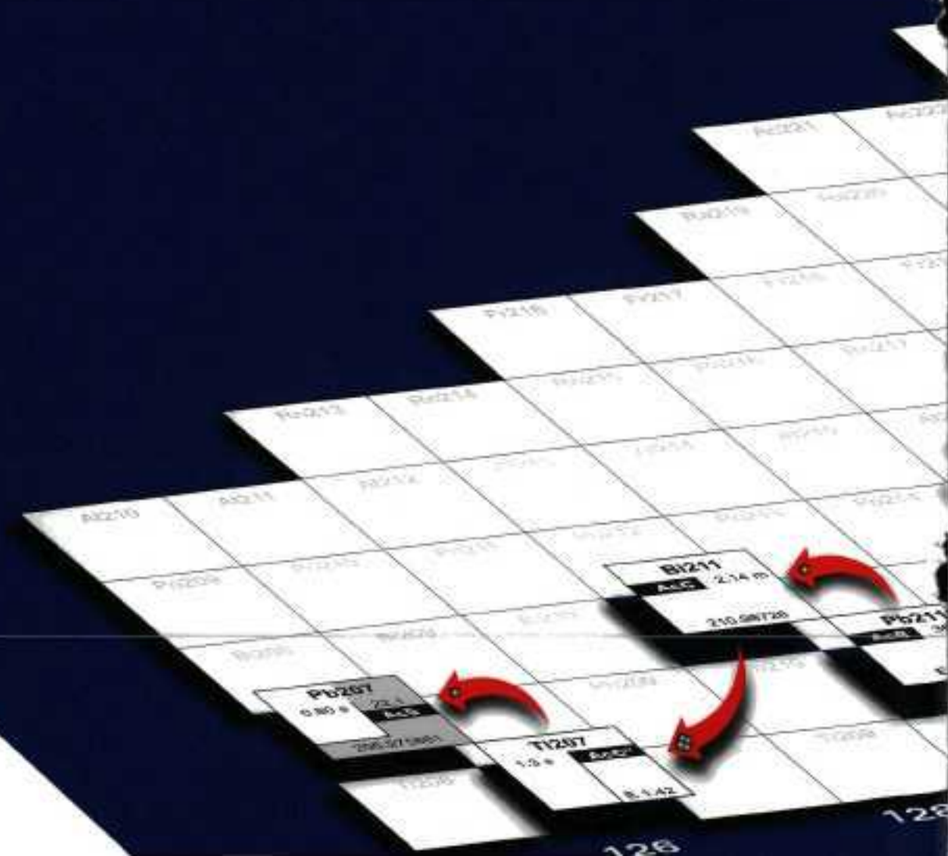

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Chart of the Nuclides

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* Significant differences in these abundance values have been observed in some samples. See booklet for discussion.

Time

μs	microseconds (1.0E-6s)
ms	milliseconds (1.0E-3s)
s	seconds
m	minutes
h	hours
d	days
a	years

Radiation and Decay

α	alpha particle
β^-	negative electron (beta decay)
β^+	positive
γ	gamma ray
n	neutron
p	proton
d	deuteron
t	triton
ε	electron capture
IT	isomeric transition
D	delayed radiation
SF	spontaneous fission
E	disintegration energy
e⁻	conversion electron
	cluster decay (Located in bottom right corner)

Other Symbols

☺	abundance less than 1%
☹	abundance less than 1.0E-3%
☠	abundance less than 1.0E-6%
E	indicates exponential format, e.g., 1.06E11a is 1.06×10^{11} years
⊸	indicates assignment to metastable (m) and ground (g) states inconclusive
⚠	indicates assignment of m₁ and m₂ states inconclusive
Z	unspecified number of particles of a given type emitted, e.g., 2P
◢	nuclide formed as a product in fission of U235, U233 or Pu239
■	shielded nuclide with independent fission yield greater than 2.5E-6%

Alphabetical List of the Elements

Name	Sym	Z	Name	Sym	Z	Name	Sym	Z
actinium	Ac	89	gold	Au	79	praseodymium	Pr	59
aluminum	Al	13	hafnium	Hf	72	promethium	Pm	61
americium	Am	95	hassium	Hs	108	protactinium	Pa	91
antimony	Sb	51	helium	He	2	radium	Ra	88
argon	Ar	18	holmium	Ho	67	radon	Rn	86
arsenic	As	33	hydrogen	H	1	rhenium	Re	75
astatine	At	85	indium	In	49	rhodium	Rh	45
barium	Ba	56	iodine	I	53	rubidium	Rb	37
berkelium	Bk	97	iridium	Ir	77	ruthenium	Ru	44
beryllium	Be	4	iron	Fe	26	rutherfordium	Rf	104
bismuth	Bi	83	krypton	Kr	36	samarium	Sm	62
bohrium	Bh	107	lanthanum	La	57	scandium	Sc	21
boron	B	5	lawrencium	Lr	103	seaborgium	Sg	106
bromine	Br	35	lead	Pb	82	selenium	Se	34
cadmium	Cd	48	lithium	Li	3	silicon	Si	14
calcium	Ca	20	lutetium	Lu	71	silver	Ag	47
californium	Cf	98	magnesium	Mg	12	sodium	Na	11
carbon	C	6	manganese	Mn	25	strontium	Sr	38
cerium	Ce	58	meitnerium	Mt	109	sulfur	S	16
cesium	Cs	55	mendelevium	Md	101	tantalum	Ta	73
chlorine	Cl	17	mercury	Hg	80	technetium	Tc	43
chromium	Cr	24	molybdenum	Mo	42	tellurium	Te	52
cobalt	Co	27	neodymium	Nd	60	terbium	Tb	65
copper	Cu	29	neon	Ne	10	thallium	Tl	81
curium	Cm	96	neptunium	Np	93	thorium	Th	90
darmstadtium	Ds	110	nickel	Ni	28	thulium	Tm	69
dubnium	Db	105	niobium	Nb	41	tin	Sn	50
dysprosium	Dy	66	nitrogen	N	7	titanium	Ti	22
einsteinium	Es	99	nobelium	No	102	tungsten	W	74
erbium	Er	68	osmium	Os	76	uranium	U	92
europium	Eu	63	oxygen	O	8	vanadium	V	23
fermium	Fm	100	palladium	Pd	46	xenon	Xe	54
fluorine	F	9	phosphorus	P	15	ytterbium	Yb	70
francium	Fr	87	platinum	Pt	78	yttrium	Y	39
gadolinium	Gd	64	plutonium	Pu	94	zinc	Zn	30
gallium	Ga	31	polonium	Po	84	zirconium	Zr	40
germanium	Ge	32	potassium	K	19			