Isotopes & Allotropy

(Important Properties of Chemical Elements)

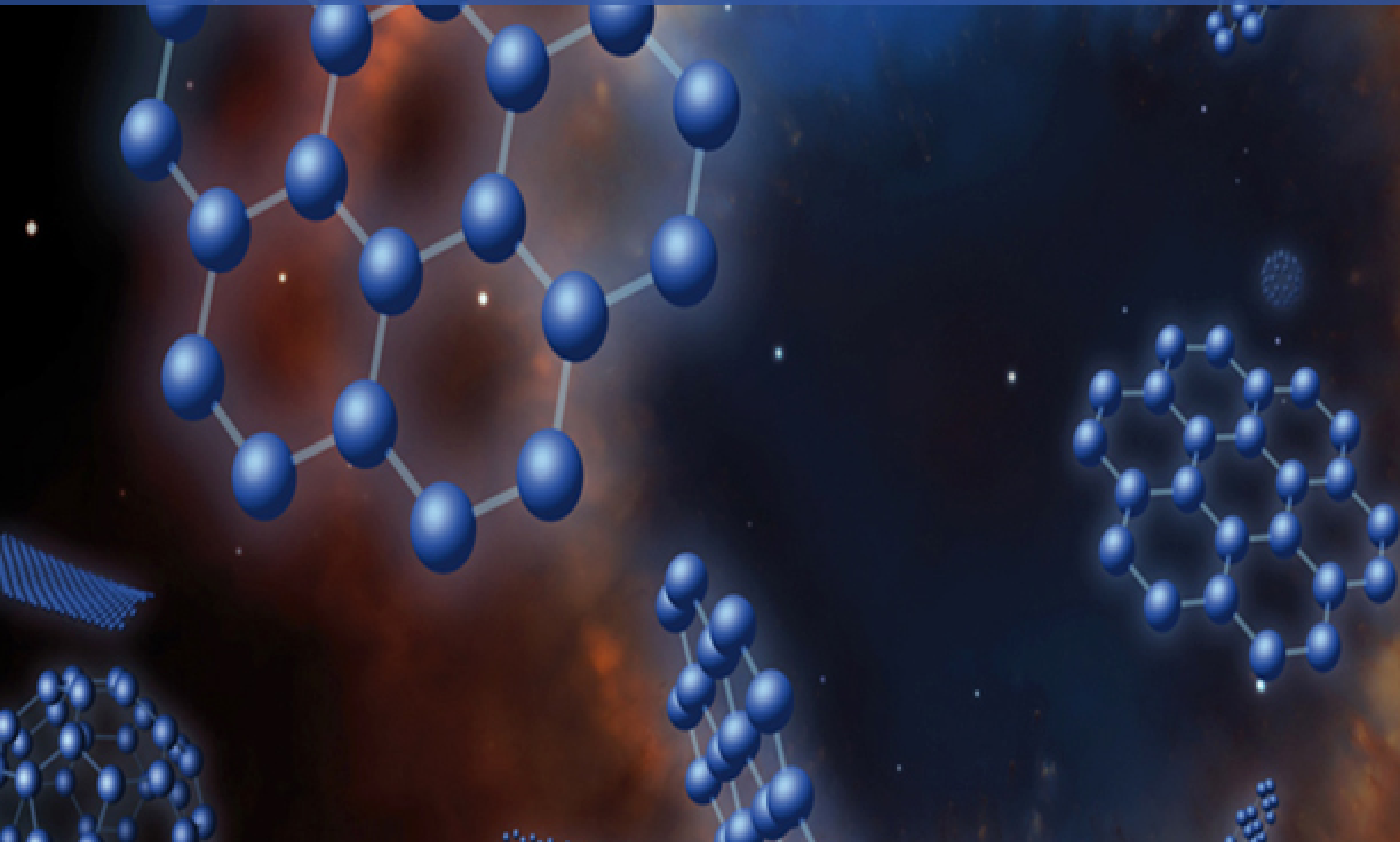

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Table of Contents

Chapter 1

Isotope

Isotopes are atoms that contain the same number of protons but a different number of neutrons. The number of protons (the atomic number) is the same for each isotope, e.g. carbon-12, carbon-13 and carbon-14 each have 6 protons, but the number of neutrons in each isotope differs. This alters the total number of nucleons (protons and neutrons) in the nucleus, known as the mass number, as well as the atomic mass. For example, carbon-12, carbon-13 and carbon-14 are three isotopes of the element carbon with mass numbers 12, 13 and 14 respectively. The atomic number of carbon is 6 (every carbon atom has 6 protons); therefore the neutron numbers in these isotopes are 6, 7 and 8 respectively.

A *nuclide* is an atom with a specific number of protons and neutrons in the nucleus, for example carbon-13 with 6 protons and 7 neutrons. The *nuclide* concept (referring to individual nuclear species) emphasizes nuclear properties over chemical properties, while the *isotope* concept (grouping all atoms of each element) emphasizes chemical over nuclear. The neutron number has drastic effects on nuclear properties, but its effect on chemical properties is negligible in most elements, and still quite small in the case of the very lightest elements, where it does matter slightly. Since *isotope* is the older term, it is better known, and is still sometimes used in contexts where *nuclide* might be more appropriate, such as nuclear technology and nuclear medicine.

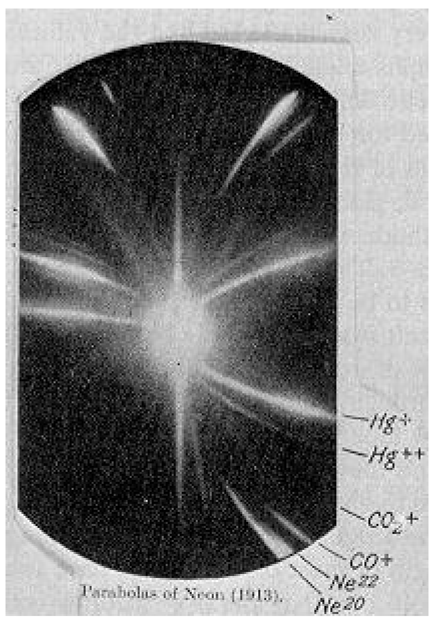
An isotope and/or nuclide is specified by the name of the particular element (this indicates the atomic number implicitly) followed by a hyphen and the mass number (e.g. helium-3, carbon-12, carbon-13, iodine-131 and uranium-238). When a chemical symbol is used, e.g., "C" for carbon, standard notation is to indicate the number of nucleons with a superscript at the upper left of the chemical symbol and to indicate the atomic number with a subscript at the lower left (e.g. 3/2He, 4/2He, 12/6C, 14/6C, 235/92U, and 239/92U). Since the atomic number is implied by the element symbol, it is common to state only the mass number in the superscript and leave out the atomic number subscript.

Some isotopes are radioactive and are therefore described as radioisotopes or radionuclides, while others have never been observed to undergo radioactive decay and are described as stable isotopes. For example, 14C is a radioactive form of carbon while12C and 13C are stable isotopes. There are about 339 naturally occurring nuclides on Earth, of which 288 are primordial nuclides, meaning that they have existed since the solar system's formation. These include 33 nuclides with very long half lives (over 80 million years) and 255 which are formally considered as "stable isotopes", since they have not been observed to decay.

Many apparently "stable" isotopes are predicted by theory to be radioactive, with extremely long half-lives (this does not count the possibility of proton decay, which would make all nuclides unstable). Of the 255 nuclides never observed to decay, only 90 of these (all from the first 40 elements) are stable in theory to all known forms of decay. Element 41 (niobium) is theoretically unstable to spontaneous fission, but this has never been detected. Many other stable nuclides are in theory energetically susceptible to other known forms of decay such as alpha decay or double beta decay, but no decay has yet been observed. The half lives for these processes often exceed a million times the estimated age of the universe, and in fact there are 27 known radionuclides with half lives longer than the age of the universe.

Adding in the radioactive nuclides that have been created artificially, there are more than 3100 currently known nuclides. These include 905 nuclides which are either stable, or have half lives longer than 60 minutes.

History



In the bottom right corner of JJ Thomson's photographic plate are the separate impact marks for the two isotopes of neon: neon-20 and neon-22.

The existence of isotopes was first suggested in 1913 by the radiochemist Frederick Soddy, based on studies of radioactive decay chains which indicated about 40 different species between uranium and lead. Since the periodic table only allows for 11 elements from uranium to lead, Soddy proposed that several types of atoms (differing in radioactive properties) can occupy the same place in the table.

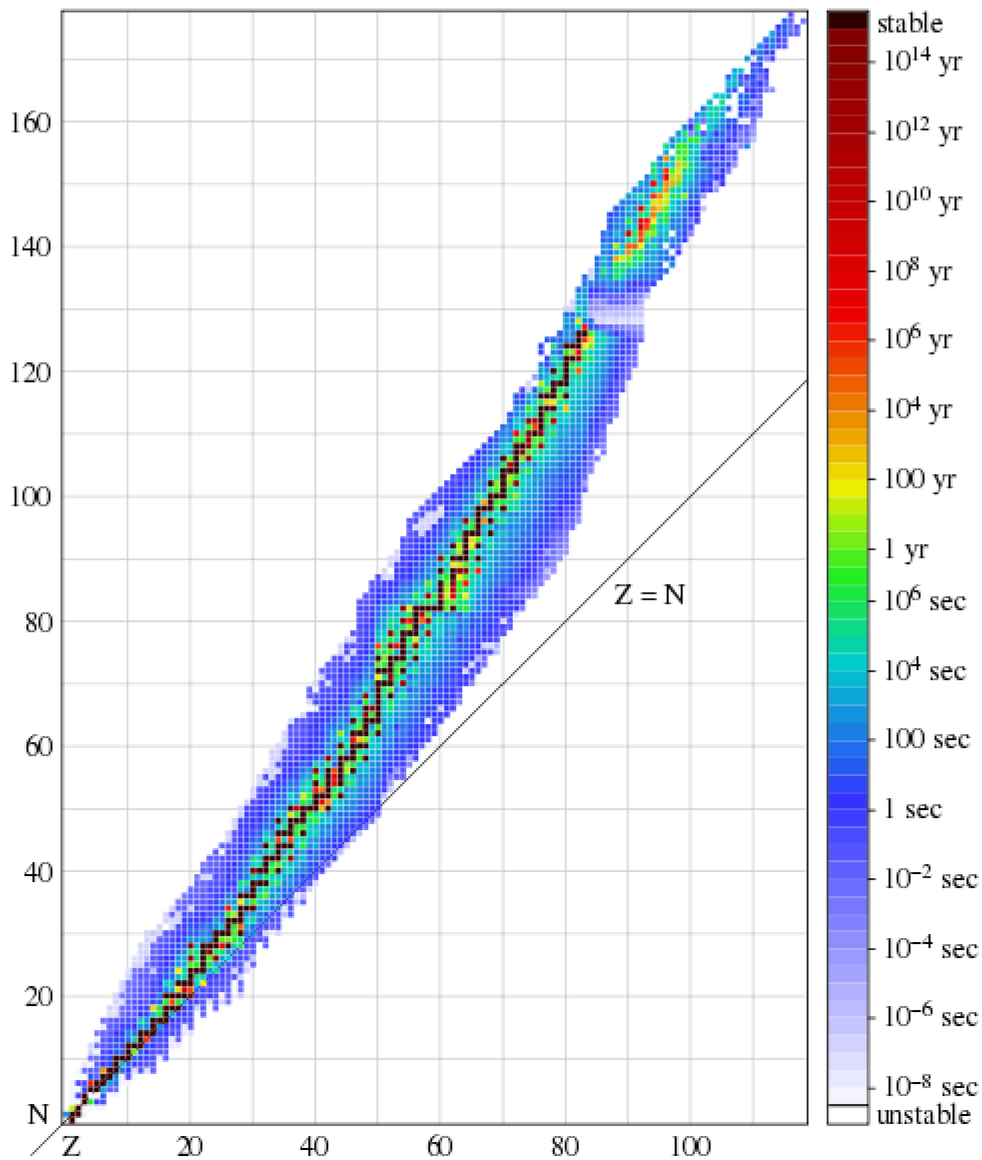
The term “isotope”, Greek for “at the same place”, was suggested to Soddy in 1914 by Margaret Todd, a Scottish physician to whom he was distantly related by marriage, during a conversation in which he explained his ideas to her.

Confirmation was provided by the observation of isotopes differing in mass for a stable (non-radioactive) element by J. J. Thomson in 1913. As part of his exploration into the composition of canal rays (positive ions), Thomson channeled streams of neon ions through a magnetic and an electric field and measured their deflection by placing a photographic plate in their path. Each stream created a glowing patch on the plate at the point it struck. Thomson observed two separate patches of light on the photographic plate, which suggested two different parabolas of deflection. Thomson eventually concluded that some of the atoms in the neon gas were of higher mass than the rest. F.W. Aston subsequently discovered different stable isotopes for numerous elements using a mass spectrograph.

Variation in properties between isotopes

Chemical and molecular properties

A neutral atom has the same number of electrons as protons. Thus, different isotopes of a given element all have the same number of protons and electrons and share a similar electronic structure. Because the chemical behavior of an atom is largely determined by its electronic structure, different isotopes exhibit nearly identical chemical behavior. The main exception to this is the kinetic isotope effect: due to their larger masses, heavier isotopes tend to react somewhat more slowly than lighter isotopes of the same element. This is most pronounced for protium (^{1}H) and deuterium (^{2}H), because deuterium has twice the mass of protium. The mass effect between deuterium and the relatively light protium also affects the behavior of their respective chemical bonds, by means of changing the center of gravity (reduced mass) of the atomic systems. However, for heavier elements, which have more neutrons than lighter elements, the ratio of the nuclear mass to the collective electronic mass is far greater, and the relative mass difference between isotopes is much less. For these two reasons, the mass-difference effects on chemistry are usually negligible.



Isotope half lives. Note that the plot for stable isotopes diverges from the line, protons Z = neutrons N as the element number Z becomes larger

In similar manner, two molecules that differ only in the isotopic nature of their atoms (*isotopologues*) will have identical electronic structure and therefore almost indistinguishable physical and chemical properties (again with deuterium providing the primary exception to this rule). The **vibrational modes** of a molecule are determined by its shape and by the masses of its constituent atoms. As a consequence, isotopologues will have different sets of vibrational modes. Since vibrational modes allow a molecule to absorb photons of corresponding energies, isotopologues have different optical properties in the infrared range.

Nuclear properties and stability

Atomic nuclei consist of protons and neutrons bound together by the residual strong force. Because protons are positively charged, they repel each other. Neutrons, which are electrically neutral, stabilize the nucleus in two ways. Their copresence pushes protons slightly apart, reducing the electrostatic repulsion between the protons, and they exert the attractive nuclear force on each other and on protons. For this reason, one or more neutrons are necessary for two or more protons to be bound into a nucleus. As the number of protons increases, so does the ratio of neutrons to protons necessary to ensure a stable nucleus. For example, although the neutron:proton ratio of $^{3}_{2}$He is 1:2, the neutron:proton ratio of $^{238}_{92}$U is greater than 3:2. A number of lighter elements have stable nuclides with the ratio 1:1 (Z = N). The nuclide $^{40}_{20}$Ca (calcium-40) is the heaviest stable nuclide with the same number of neutrons and protons; all heavier stable nuclides contain more neutrons than protons.

Numbers of isotopes per element

Of the 80 elements with a stable isotope, the largest number of stable isotopes observed for any element is ten (for the element tin). Xenon is the only element that has nine stable isotopes. No element has eight stable isotopes. Four elements have seven stable isotopes, nine have six stable isotopes, nine have five stable isotopes, nine have four stable isotopes, five have three stable isotopes, 16 have two stable isotopes (counting $^{180m}_{73}$Ta as stable), and 26 elements have only a single stable isotope (of these, 19 are so-called mononuclidic elements, having a single primordial stable isotope that dominates and fixes the atomic weight of the natural element to high precision; 3 radioactive **mononuclidic** elements occur as well). In total, there are 255 nuclides that have not been observed to decay. For the 80 elements that have one or more stable isotopes, the average number of stable isotopes is 255/80 = 3.2 isotopes per element.

Even and odd nucleon numbers

Even/odd N

Mass	**E**	**O**	**All**
Stable	145	101	246
Long-lived	20	6	26
Primordial	165	107	272

The proton:neutron ratio is not the only factor affecting nuclear stability. Adding neutrons to isotopes can vary their nuclear spins and nuclear shapes, causing differences in neutron capture cross-sections and gamma spectroscopy and nuclear magnetic resonance properties.

Even mass number

Beta decay of an even-even nucleus produces an odd-odd nucleus, and vice versa. An even number of protons or of neutrons are more stable (lower binding energy) because of pairing effects, so even-even nuclei are much more stable than odd-odd. One effect is that there are few stable odd-odd nuclei, but another effect is to prevent beta decay of many even-even nuclei into another even-even nucleus of the same mass number but lower energy, because decay proceeding one step at a time would have to pass through an odd-odd nucleus of higher energy. This makes for a larger number of stable even-even nuclei, up to three for some mass numbers, and up to seven for some atomic (proton) numbers. Double beta decay directly from even-even to even-even skipping over an odd-odd nuclide is only occasionally possible, and even then with a half-life greater than a billion times the age of the universe. For example, the double beta emitter ^{116}Cd has a half-life of 2.9×10^{19} years.

Even-mass-number nuclides have integer spin and are bosons.

Even proton-even neutron

Even/odd Z, N

p,n	**EE**	**OO**	**EO**	**OE**
Stable	140	5	53	48
Long-lived	16	4	2	4
Primordial	156	9	55	52

For example, the extreme stability of helium-4 due to a double pairing of 2 protons and 2 neutrons prevents *any* nuclides containing five or eight nucleons from existing for long enough to serve as platforms for the buildup of heavier elements during fusion formation in stars.

There are 141 stable even-even isotopes, forming 55% of the 255 stable isotopes. There are also 16 primordial long-lived even-even isotopes. As a result, many of the 41 even-numbered elements from 2 to 82 have many primordial isotopes. Half of these even-numbered elements have six or more stable isotopes.

All even-even nuclides have spin 0 in their ground state.

Odd proton-odd neutron

Only five stable nuclides contain both an odd number of protons *and* an odd number of neutrons: the first four odd-odd nuclides $^{2}_{1}$H, $^{6}_{3}$Li, $^{10}_{5}$B, and $^{14}_{7}$N (where changing a proton to a neutron or vice versa would lead to a very lopsided proton-neutron ratio) and $^{180m}_{73}$Ta, which has not yet been observed to decay despite experimental attempts. Also, four long-lived radioactive odd-odd nuclides ($^{40}_{19}$K, $^{50}_{23}$V, $^{138}_{57}$La, $^{176}_{71}$Lu) occur naturally.

Of these 9 primordial odd-odd nuclides, only $^{14}_{7}$N is the most common isotope of a common element, because it is a part of the CNO cycle; $^{6}_{3}$Li and $^{10}_{5}$B are minority isotopes of elements that are rare compared to other light elements, while the other six isotopes make up only a tiny percentage of their elements.

Few odd-odd nuclides (and none of the primordial ones) have spin 0 in the ground state.

Odd mass number

There is only one beta-stable nuclide per odd mass number because there is no difference in binding energy between even-odd and odd-even comparable to that between even-even and odd-odd, and other nuclides of the same mass are free to beta decay towards the lowest-energy one. For mass numbers 5, 147, 151, and 209 and up, the one beta-stable isobar is able to alpha decay, so that there are no stable isotopes with these mass numbers. This gives a total of 101 stable isotopes with odd mass numbers.

Odd-mass-number nuclides have half-integer spin and are fermions.

Odd proton-even neutron

These form most of the stable isotopes of the odd-numbered elements, but there is only one stable odd-even isotope for each of the 41 odd-numbered elements from 1 to 81, except for technetium (43Tc) and promethium (61Pm) that have no stable isotopes, and chlorine (17Cl), potassium (19K), copper (29Cu), gallium (31Ga), bromine (35Br), silver (47Ag), antimony (51Sb), iridium (Ir), and thallium (81Tl), each of which has two, making a total of 48 stable odd-even isotopes. There are also four primordial long-lived odd-even isotopes, 87 37Rb, 115 49In, 151 63Eu, and 187 75Re.

Even proton-odd neutron

There are 54 stable isotopes that have an even number of protons and an odd number of neutrons. There are also four primordial long lived even-odd isotopes, 113 48Cd (beta decay, half-life is 7.7×10^{15} years); 147 62Sm (1.06×10^{11}a); and 149 62Sm ($>2\times10^{15}$a); and the fissile 235 92U.

The only even-odd isotopes that are the most common one for their element are 195 78Pt and 9 4Be. Beryllium-9 is the only stable beryllium isotope because the expected beryllium-8 has higher energy than two alpha particles and therefore decays to them.

Odd neutron number

Even/odd N

n	E	O
Stable	188	58
Long-lived	20	6
Primordial	208	64

The only odd-neutron-number isotopes that are the most common isotope of their element are 195 78Pt, 9 4Be and 14 7N.

Actinides with odd neutron number are generally fissile, while those with even neutron number are generally not, though they are split when bombarded with fast neutrons.

Occurrence in nature

Elements are composed of one or more naturally occurring isotopes. The unstable (radioactive) isotopes are either primordial, in which case they have persisted down to the present because their rate of decay is so slow (e.g., uranium-238 and potassium-40), or they are postprimordial, created by cosmic ray bombardment as cosmogenic nuclides (e.g., tritium, carbon-14) or by the decay of a radioactive primordial isotope to a radioactive radiogenic nuclide daughter (e.g., uranium to radium).

As discussed above, only 80 elements have any stable isotopes, and 26 of these have only one stable isotope. Thus, about two thirds of stable elements occur naturally on Earth in multiple stable isotopes, with the largest number of stable isotopes for an element being ten, for tin ($_{50}Sn$). There are about 94 elements found naturally on Earth (up to plutonium inclusive), though some are detected only in very tiny amounts, such as plutonium-244. Scientists estimate that the elements that occur naturally on Earth (some only as radioisotopes) occur as 339 isotopes (nuclides) in total. Only 255 of these naturally occurring isotopes are stable in the sense of never having been observed to decay as of the present time An additional 33 primordial nuclides (to a total of 288 primordial nuclides), are radioactive with known half lives, but have half lives longer than 80 million years, allowing them to exist from the beginning of the solar system.

All the known stable isotopes occur naturally on Earth; the other naturally occurring-isotopes are radioactive but occur on Earth due to their relatively long half-lives, or else due to other means of ongoing natural production. These include the afore-mentioned cosmogenic nuclides and the short-lived radioisotopes formed by decay of a primordial radioactive isotope, such as radon and radium from uranium.

An additional ~3000 radioactive isotopes not found in nature have been created in nuclear reactors and in particle accelerators. Many short-lived isotopes not found naturally on Earth have also been observed by spectroscopic analysis, being naturally created in stars or supernovae. An example is aluminum-26, which is not naturally found on Earth, but which is found in abundance on an astronomical scale.

The tabulated atomic masses of elements are averages that account for the presence of multiple isotopes with different masses. Before the discovery of isotopes, empirically determined noninteger values of atomic mass confounded scientists. For example, a sample of chlorine contains 75.8% chlorine-35 and 24.2% chlorine-37, giving an average atomic mass of 35.5 atomic mass units.

According to generally accepted cosmology theory, only isotopes of hydrogen and helium, traces of some isotopes of lithium and beryllium, and perhaps some boron, were created at the Big Bang, while all other isotopes were synthesized later, in stars and supernovae, and in interactions between energetic particles such as cosmic rays, and previously produced isotopes. The respective abundances of isotopes on Earth result from the quantities formed by these processes, their spread through the galaxy, and the rates of decay for isotopes that are unstable. After the initial coalescence of the solar system, isotopes were redistributed according to mass, and the isotopic composition of elements varies slightly from planet to planet. This sometimes makes it possible to trace the origin of meteorites.

Atomic mass of isotopes

The atomic mass (m_r) of an isotope is determined mainly by its mass number (i.e. number of nucleons in its nucleus). Small corrections are due to the binding energy of the nucleus, the slight difference in mass between proton and neutron, and the mass of the electrons associated with the atom, the latter because the electron:nucleon ratio differs among isotopes.

The mass number is a dimensionless quantity. The atomic mass, on the other hand, is measured using the atomic mass unit based on the mass of the carbon atom. It is denoted with symbols "u" (for unit) or "Da" (for Dalton).

The atomic masses of naturally occurring isotopes of an element determine the atomic weight of the element. When the element contains *N* isotopes, the equation below is applied for the atomic weight *M*:

$$M = m_1 x_1 + m_2 x_2 + ... + m_N x_N$$

where m_1, m_2, ..., m_N are the atomic masses of each individual isotope, and x_1, ... , x_N are the relative abundances of these isotopes.

Applications of isotopes

Several applications exist that capitalize on properties of the various isotopes of a given element. Isotope separation is a significant technological challenge, particularly with heavy elements such as uranium or plutonium. Lighter elements such as lithium, carbon, nitrogen, and oxygen are commonly separated by gas diffusion of their compounds such as CO and NO. The separation of hydrogen and deuterium is unusual since it is based on chemical rather than physical properties, for example in the Girdler sulfide process. Uranium isotopes have been separated in bulk by gas diffusion, gas centrifugation, laser ionization separation, and (in the Manhattan Project) by a type of production mass spectrometry.

Use of chemical and biological properties

- Isotope analysis is the determination of isotopic signature, the relative abundances of isotopes of a given element in a particular sample. For biogenic substances in particular, significant variations of isotopes of C, N and O can occur. Analysis of such variations has a wide range of applications, such as the detection of adulteration of food products. The identification of certain meteorites as having originated on Mars is based in part upon the isotopic signature of trace gases contained in them.

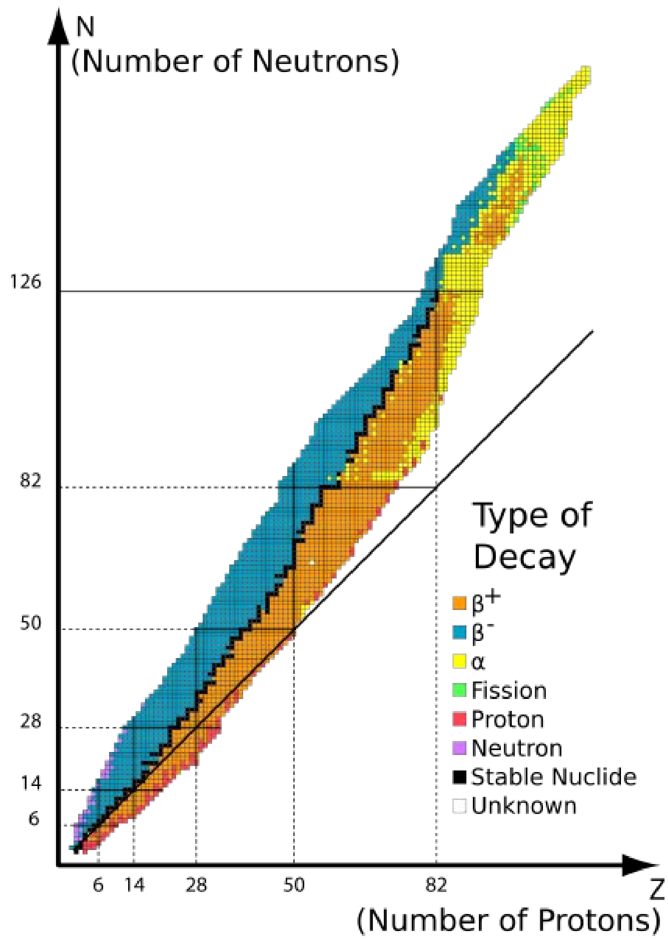
- Another common application is isotopic labeling, the use of unusual isotopes as tracers or markers in chemical reactions. Normally, atoms of a given element are indistinguishable from each other. However, by using isotopes of different masses, they can be distinguished by mass spectrometry or infrared spectroscopy. For example, in 'stable isotope labeling with amino acids in cell culture (SILAC)' stable isotopes are used to quantify proteins. If radioactive isotopes are used, they can be detected by the radiation they emit (this is called *radioisotopic labeling*).
- A technique similar to radioisotopic labeling is radiometric dating: using the known half-life of an unstable element, one can calculate the amount of time that has elapsed since a known level of isotope existed. The most widely known example is radiocarbon dating used to determine the age of carbonaceous materials.
- Isotopic substitution can be used to determine the mechanism of a reaction via the kinetic isotope effect.

Use of nuclear properties

- Several forms of spectroscopy rely on the unique nuclear properties of specific isotopes. For example, nuclear magnetic resonance (NMR) spectroscopy can be used only for isotopes with a nonzero nuclear spin. The most common isotopes used with NMR spectroscopy are ^{1}H, ^{2}D, ^{15}N, ^{13}C, and ^{31}P.
- Mössbauer spectroscopy also relies on the nuclear transitions of specific isotopes, such as ^{57}Fe.
- Radionuclides also have important uses. Nuclear power and nuclear weapons development require relatively large quantities of specific isotopes.

Chapter 2

Stable Isotope



Graph of isotopes/nuclides by type of decay. Orange and blue nuclides are unstable, with the black squares between these regions representing stable nuclides. The unbroken line passing below many of the nuclides represents the theoretical position on the graph of nuclides for which proton number is the same as neutron number. The graph shows that elements with more than 20 protons must have more neutrons than protons, in order to be stable.

Stable isotopes are chemical isotopes that may or may not be radioactive, but if radioactive, have half lives too long to be measured.

Only 90 nuclides from the first 40 elements are energetically stable to any kind of decay save proton decay, in theory. An additional 165 are theoretically unstable to known types of decay, but no evidence of decay has ever been observed, for a total of 255 nuclides for which there is no evidence of radioactivity. By this definition, there are 255 known stable nuclides of the 80 elements which have one or more stable isotopes.

Of the 80 elements with one or more stable isotopes, twenty-six have only a single stable isotope, and are thus termed *monoisotopic*, and the rest have more than one stable isotope. One element (tin) has ten stable isotopes, the largest number known for an element.

Properties of stable isotopes

Different isotopes of the same element (whether stable or unstable) have nearly the same chemical characteristics and therefore behave almost identically in biology (a notable exception is the isotopes of hydrogen). The mass differences, due to a difference in the number of neutrons, will result in partial separation of the light isotopes from the heavy isotopes during chemical reactions and during physical processes such as diffusion and vaporization. This process is called isotope fractionation. For example, the difference in mass between the two stable isotopes of hydrogen, ^{1}H (1 proton, no neutron, also known as protium) and ^{2}H (1 proton, 1 neutron, also known as deuterium) is almost 100%. Therefore, a significant fractionation will occur.

Study of stable isotopes

Commonly analysed stable isotopes include oxygen, carbon, nitrogen, hydrogen and sulfur. These isotope systems have been under investigation for many years in order to study processes of isotope fractionation in natural systems because they are relatively simple to measure. Recent advances in mass spectrometry (i.e. multiple-collector inductively coupled plasma mass spectrometry) now enable the measurement of heavier stable isotopes, such as iron, copper, zinc, molybdenum, etc.

Stable isotopes have been used in botanical and plant biological investigations for many years, and more and more ecological and biological studies are finding stable isotopes (mostly carbon, nitrogen and oxygen) to be extremely useful. Other workers have used oxygen isotopes to reconstruct historical atmospheric temperatures, making them important tools for climate research. Measurements of ratios of one naturally occurring stable isotope to another play an important role in radiometric dating and isotope geochemistry, and also helpful for determining patterns of rainfall and movements of elements through living organisms, helping sort out food web dynamics in ecosystems.

Definition of stability, and natural isotopic presence

Most naturally occurring nuclides are stable and about 33 more (total of 288) are known radioactives with sufficiently long half-lives (also known) to occur "primordially." If the half-life of a nuclide is comparable to, or greater than, the Earth's age (4.5 billion years), a significant amount will have survived since the formation of the Solar System, and then is said to be primordial. It will then contribute in that way to the natural isotopic composition of a chemical element. Primordially present radioisotopes are easily detected with half-lives as short as 700 million years (e.g., ^{235}U), although some primordial isotopes have been detected with half lives as short as 80 million years (e.g., ^{244}Pu). However, this is the present limit of detection, as the nuclide with the next-shortest half life (niobium-92 with half life 34.7 million years) has not been yet been detected in nature.

Many naturally-occurring radioisotopes (another 51 or so, for a total of about 339) exhibit still shorter half-lives than 80 million years, but they are made freshly, as daughter products of decay processes of primordial nuclides (for example, radium from uranium) or from ongoing energetic reactions, such as cosmogenic nuclides produced by present bombardment of Earth by cosmic rays (for example, carbon-14 made from nitrogen).

Many isotopes that are classed as stable (i.e. no radioactivity has been observed for them) are predicted to have extremely long half-lives (sometimes as high as 10^{18} years or more). If the predicted half-life falls into an experimentally accessible range, such isotopes have a chance to move from the list of stable nuclides to the radioactive category, once their activity is observed. Good examples are bismuth-209 and tungsten-180 which were formerly classed as stable, but have been recently (2003) found to be alpha-active. However, such nuclides do not change their status as primordial when they are found to be radioactive.

Most stable isotopes in the earth are believed to have been formed in processes of nucleosynthesis, either in the 'Big Bang', or in generations of stars that preceded the formation of the solar system. However, some stable isotopes also show abundance variations in the earth as a result of decay from long-lived radioactive nuclides. These decay-products are termed radiogenic isotopes, in order to distinguish them from the much larger group of 'non-radiogenic' isotopes.

Research areas

The so-called Island of Stability may reveal a number of long-lived or even stable atoms that are heavier (and with more protons) than lead.

Stable isotope fractionation

There are three types of isotope fractionation:

- equilibrium fractionation
- kinetic fractionation
- mass-independent fractionation

Isotopes per element

Of the known chemical elements, 80 elements have at least one stable nuclide. These comprise the first 82 elements from hydrogen to lead, with the exceptions of technetium (#43) and promethium (#61), which do not have any stable nuclides. As of December, 2010, there were a total of 255 known "stable" nuclides. In this definition, "stable" means a nuclide which has either never been observed to decay against the natural background. Thus, these elements have half lives too long to be measured by any means, direct or indirect.

Only one element (tin) has 10 stable isotopes, and one (xenon) has nine stable isotopes. No elements have exactly eight stable isotopes, but four elements have seven stable isotopes, nine have six stable isotopes, nine have five stable isotopes, nine have four stable isotopes, five have three stable isotopes, 16 have two stable isotopes, and 26 have only a single stable isotope and are thus considered **monoisotopic** elements. The mean number of stable isotopes for elements which have at least one such isotope, is 255/80 = 3.2.

"Magic numbers" and odd and even proton and neutron count

Stability of isotopes is affected by the ratio of protons to neutrons, and also by presence of certain "magic numbers" of neutrons or protons which represent closed and filled quantum shells. These quantum shells correspond to a set of energy levels within the shell model of the nucleus; filled shells, such as the filled shell of 50 protons for tin, confers unusual stability on the nuclide. As in the case of tin, a magic number for **Z**, the atomic number, tends to increase the number of stable isotopes for the element.

Just as in the case of electrons, which have the lowest energy state when they occur in pairs in a given orbital, nucleons (both protons and neutrons) exhibit a lower energy state when their number is even, rather than odd. This stability tends to prevent beta decay (in two steps) of many even-even nuclides into another even-even nuclide of the same mass number but lower energy (and of course with two more protons and two fewer neutrons), because decay proceeding one step at a time would have to pass through an odd-odd nuclide of higher energy. This makes for a larger number of stable even-even nuclides, up to three for some mass numbers, and up to seven for some atomic (proton) numbers. Conversely, of the 255 known stable nuclides, only four have both an odd number of protons *and* odd number of neutrons: hydrogen-2 (deuterium), lithium-6, boron-10 and nitrogen-14. Also, only four naturally occurring, radioactive odd-odd nuclides have a half-life over a billion years: potassium-40, vanadium-50, lanthanum-138 and tantalum-180m. Odd-odd primordial nuclides are rare because most odd-odd nuclei are highly unstable with respect to beta decay, because the decay products are even-even, and are therefore more strongly bound, due to nuclear pairing effects.

Yet another effect of the instability of an odd number of either type of nucleons, is that odd-numbered elements tend to have fewer stable isotopes. Of the 26 monoisotopic elements that have only a single stable isotope, all but one have an odd atomic number —