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IUPAC Periodic Table of Isotopes for the Educational Community

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IUPAC Periodic Table of Isotopes for the Educational Community

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Foreword—Here we report on progress of IUPAC project 2007-038-3-200 (<http://www.iupac.org/web/ins/2007-038-3-200>) titled “Development of an Isotopic Periodic Table for the Educational Community,” whose other members include

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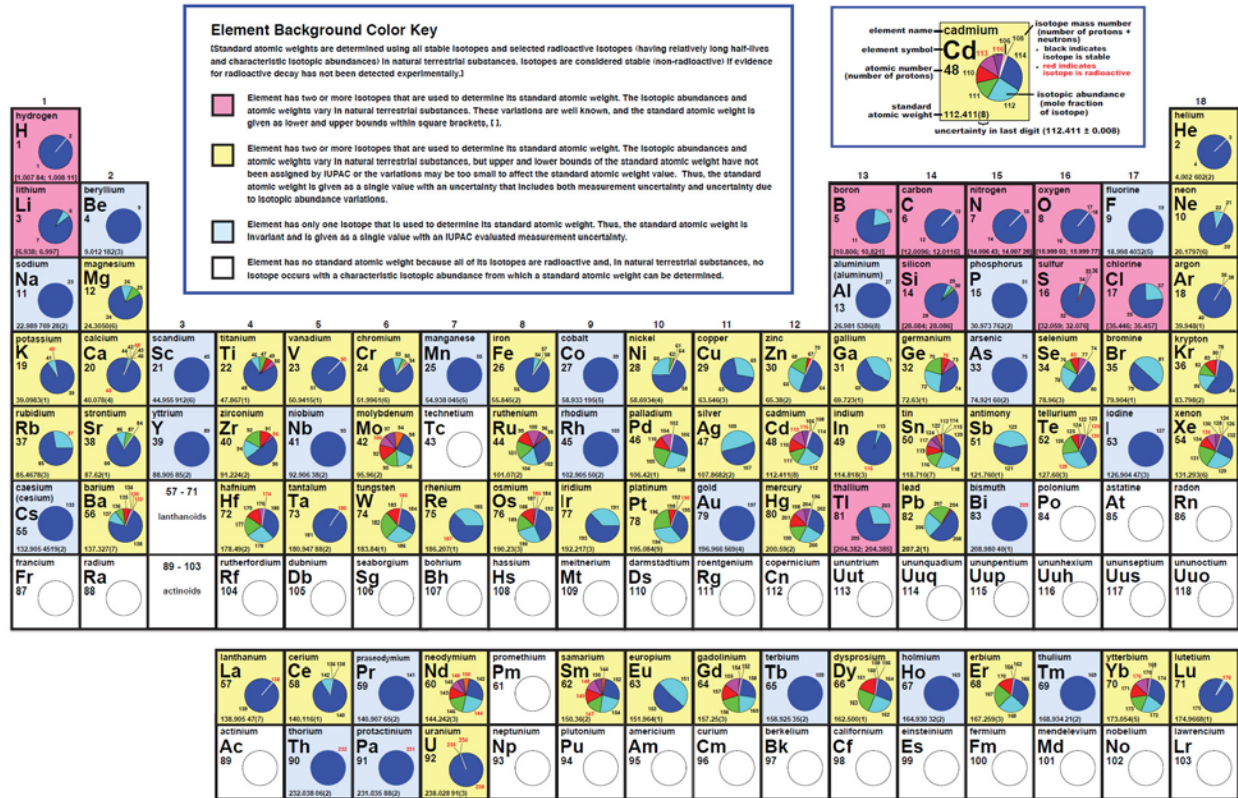
More than two person-years have gone into the development of this Periodic Table of the Isotopes (Figure 1), which we envision may reside on walls of chemistry laboratories near the Periodic Table of the Elements.

Introduction

John Dalton first proposed the concept of atomic weights of the elements in the first decade of the nineteenth century. These atomic weights of the chemical elements were thought of as constants of nature, similar to the speed of light. Dmitri Mendeleev arranged the atomic weights of the elements in ascending order of value and used the systematic variation of their chemical properties to produce his Periodic Table of the Elements in 1869. Measurement of atomic weight values became an important chemical activity for a century and a half. Theodore Richards received a Noble Prize for his work in this area.

In 1913, Fredrick Soddy found a species of radium, which had an atomic weight value of 228, compared to the familiar radium gas value of 226. Soddy coined the term “isotope” (Greek for “in the same place”) to account for this second atomic weight value in the radium position of the Periodic Table. Both of these isotopes of radium are radioactive. Radioactive isotopes are energetically unstable and will decay (disintegrate) over time. The time it takes for one half of a sample of a given radioactive isotope to decay is the half-life of that isotope. In addition to having different atomic weight values, radium-226 and radium-228 also have different half-life values. Around the same time as Soddy’s work, J.J. Thomson (discoverer of the electron) identified two stable (non-radioactive) isotopes of the same element, neon. Over the next 40 years, the majority of the known chemical elements were found to have two or more stable (or long-lived radioactive isotopes that contribute significantly to the determination of the atomic weights of the elements).

IUPAC Periodic Table of the Isotopes



International Union of Pure and Applied Chemistry (IUPAC) Project 2007-038-3-200, "Development of an isotopic periodic table for the educational community"

May 25, 2012 www.ciaaw.org

Figure 1. IUPAC Periodic Table of the Isotopes (front side, version of May 25, 2012). As this Table is updated, it will be posted at the Web site of the Commission on Isotopic Abundances and Atomic Weights (www.ciaaw.org) and at the Web page for this IUPAC project (<http://www.iupac.org/web/ins/2007-038-3-200>).

The atomic weight (also called the relative atomic mass), $A_r(E)$, of element E in a substance P is expressed by the relation

$$A_r(E)_P = \sum [x(^iE) \times A_r(^iE)]$$

where $x(^iE)$ is the mole fraction of isotope iE (also called the isotopic abundance) and the summation is over all stable isotopes and selected radioactive isotopes (having relatively long half-lives and characteristic isotopic abundances) of the element (<http://www.ciaaw.org/pubs/TSAW%202009.pdf>). The standard atomic weight is determined from evaluation of published data by the Commission on Isotopic Abundances and Atomic Weights (CIAAW) of the International Union of Pure and Applied Chemistry (IUPAC) and is the best estimate of atomic weights of an element that might be found in natural terrestrial substances. The standard atomic weight of each element is dimensionless and is based upon assignment of a value of 12 exactly to the atomic weight of an unbound neutral carbon-12 atom in its nuclear and electronic ground state.

By the 2nd half of the twentieth century, the standard atomic weight of an element determined from the abundances and masses of isotopes provided a better estimate (a smaller measurement uncertainty) than the value determined by chemical measurement. Today, all new standard atomic weight values are determined from measurements of abundances and masses of their isotopes. For many chemical elements, it was discovered that their isotopic abundance values varied in different natural terrestrial samples. These variations in isotopic abundances led to variations in atomic weight values of an element. In the 1950s, the measurement uncertainty of standard atomic weight values was larger than the uncertainty due to natural isotopic variation for most elements. However, in the past five decades, improved measurements of isotopic abundances for many elements in different natural terrestrial samples yielded atomic weight values with low uncertainties that differed substantially from each other. Thus, standard atomic weight values could no longer be considered “constants of nature”, as had been commonly thought by some for the past two centuries. To bring attention to the fact that standard atomic weights of some elements are not constants of nature, the CIAAW introduced the concept of “intervals” (see below) to the standard atomic weight to account for the variation in isotopic abundances and atomic weight values in natural terrestrial substances (<http://www.ciaaw.org/pubs/TSAW%202009.pdf>).

Understanding Isotopes of Chemical Elements

The atoms of all elements are made up of a positively-charged nucleus surrounded by an equal amount of negative charge carried by surrounding electrons. Nuclei themselves are generally made up of electrically positive charged particles called protons and electrically neutral particles called neutrons (the lightest hydrogen isotope, not having any neutrons, being an exception). The number of protons in each atom (the atomic number, with symbol Z) determines the chemical element; for example, $Z = 1$ is hydrogen, $Z = 6$ is carbon, and $Z = 79$ is gold. The number of neutrons (symbol N) in an atom of a given element may vary. The total number of protons and neutrons ($Z + N$) in a specific atom is the mass number (symbol A ; thus, $A = Z + N$). A nuclide is an atom designated by its atomic number and its mass number, written in the form of the element name (or symbol) and mass number; for example, aluminium-27 and ^{27}Al both refer to the nuclide of the element aluminium (commonly written aluminum in the U.S.) with mass number 27. Nuclides of a given element that have different numbers of neutrons (but the same number of protons) are called isotopic nuclides or isotopes. For any particular element, only certain isotopes are stable. ^{27}Al , with 13 protons and 14 neutrons, is stable. ^{28}Al , with 15 neutrons, is unstable. A stable isotope is defined as an isotope for which no radioactive decay has been experimentally detected. Isotopes that are unstable are called radioactive isotopes (or radioisotopes). The term isotope applies equally to both stable and radioactive isotopes.

The world that surrounds us, including the water we drink and the air we breathe, is mostly made of stable isotopes of the common elements, e.g., hydrogen, oxygen, and nitrogen. With the improvement in instrumentation over time, many isotopes that were once considered stable are now known to be unstable. They undergo radioactive decay, although with very long half-lives, and they are called long-lived radioactive isotopes. In natural terrestrial substances, a radioactive isotope with a sufficiently long half-life is said to have a characteristic isotopic abundance, and it may contribute to the standard atomic weight of an element if its isotopic abundance is sufficiently large. An example is calcium-40.

Of all chemical elements that have been discovered in nature (in contrast to elements that have been synthesized or produced by man), their natural terrestrial samples contain a total of 288 different nuclides. These nuclides originally were thought to be all stable isotopes until radioactivity for some was discovered. Two hundred forty-nine of these are in fact stable isotopes, but 39 have since been determined to be long-lived radioactive isotopes that have a characteristic isotopic abundance and contribute to a standard atomic weight. In addition, more than 3,000 other nuclides are known, and they correspond to radioactive isotopes of all elements, most with short half-lives.

The Periodic Table of the Isotopes

As part of the International Year of Chemistry (IYC-2011), IUPAC funded the effort to produce the IUPAC Periodic Table of the Isotopes (IPTI), which is shown in Figure 1 and is modeled after the Periodic Table of the Elements. While the Periodic Table of the Elements indicates the similarities in the chemical properties of the elements, the IPTI emphasizes the uniqueness of each element. Each element block of the IPTI provides the chemical name, the chemical symbol, the atomic number, and the standard atomic weight of that element. A color-coded pie chart displays all of the stable isotopes and radioactive isotopes having characteristic isotopic abundances that contribute to the standard atomic weight of an element. The mole fraction (or isotopic abundance) of each of these isotopes is depicted by the relative size of the pie slice associated with that isotope. The mass number of each isotope appears around the outside of the pie chart. Each mass number is shown in black for stable isotopes and in red for radioactive isotopes.

The background color scheme used for cells of the elements on the IPTI depends in part on the number of isotopes that are used to determine the standard atomic weight of an element. If only one isotope is used to determine the standard atomic weight, the standard atomic weight is invariant and is given as a single value with an IUPAC evaluated measurement uncertainty, and the background color is blue. An example is aluminium with a standard atomic weight value of 26.981 5386(8), where the uncertainty in the last digit is indicated by the value in parentheses. If an element has no standard atomic weight because all of its isotopes are radioactive and, in natural terrestrial substances, no isotope occurs with a characteristic isotopic abundance from which a standard atomic weight can be determined, the background color is white. If an element has two or more isotopes that are used to determine its standard atomic weight and the variations in isotopic abundances and atomic weights in natural terrestrial substances are well known, the standard atomic weight is given as lower and upper bounds within square brackets, [], and the background color of the element is pink. An example is boron with a standard atomic weight of [10.806; 10.821], and this form indicates that atomic weight values are found in natural terrestrial substances as low as 10.806 and as high as 10.821 as shown in Figure 2. If an element has two or more isotopes that are used to determine its standard atomic weight, but the upper and lower bounds of the standard atomic weight have not been assigned by IUPAC or if the variations are too small to affect the standard atomic weight value, the standard atomic weight is given as a single value with an uncertainty that includes both measurement uncertainty and uncertainty due to isotopic abundance variations. The background color of such elements is yellow, and an example is osmium with a standard atomic weight of 190.23(3).

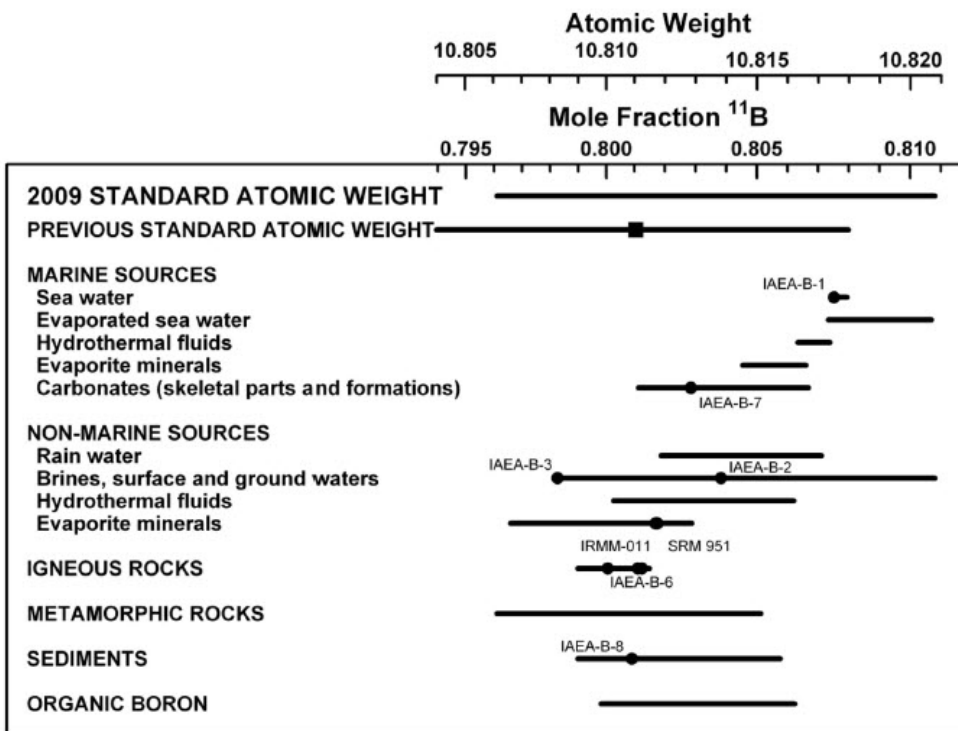


Figure 2. Variation in atomic weight with isotopic composition of selected boron-bearing materials from figure 5 of the document at <http://www.ciaaw.org/pubs/TSAW%202009.pdf>.

In a follow-on IUPAC project, an electronic version of the IPTI is planned in which a click on a chemical element cell by a user will open a form window to display additional information about the isotopes of that element. This window will include a table of the isotopes in natural terrestrial substances, their atomic masses, and their isotopic abundances. In addition, a figure will display all isotopes of that element, both stable and radioactive. The stable or naturally occurring radioactive isotopes with a characteristic isotopic abundance will be shown in the same color as they appear in the element's pie chart. Other radioactive isotopes will be depicted in one of three half-life ranges: less than one hour, greater than one year, and the intermediate range between one hour and one year. To stress the importance of isotopes in our everyday lives and why teachers, students and the general public should be aware of isotopes, some applications of selected stable and radioactive isotopes will be provided in one or more of the following general categories:

- Isotopes in industry
- Isotopes in medicine
- Isotopes in geochronology
- Isotopes in earth and planetary science applications
- Isotopes in forensic science and anthropology
- Isotopes as sources for making new isotopes
- Isotopes in biology

In each case, the application of a specified isotope will be described and often depicted pictorially. The scope of these applications will be vast and should impress readers with how often isotopes affect our daily lives.

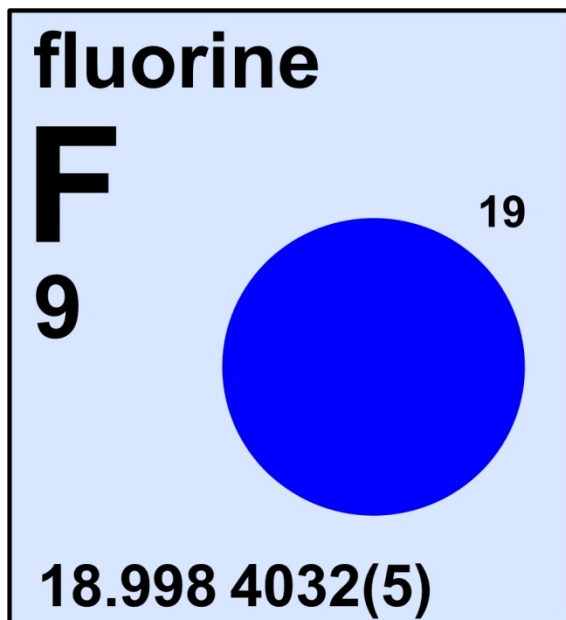
An example of the applications of isotopes of the chemical elements is displayed in appendix A for the element fluorine. In this case, there is only one isotope, fluorine-19 (^{19}F), that is stable and determines the atomic weight. The background color for the pie chart is blue and the single stable isotope of mass number 19 is shown in black. In the figure in appendix A displaying all of the isotopes of fluorine, note that only ^{19}F has its mass number displayed in black. All other isotopes have their mass numbers displayed in red because they are radioactive. All radioactive isotopes of fluorine have a half-life less than one hour, except for ^{18}F . They are shown on a white background because of their short half-life, whereas ^{18}F has a dotted background because its half-life is between one hour and one year.

An example of the application of a fluorine isotope in everyday life is shown for ^{18}F in the category of isotopes in medicine. It is used for imaging the organs, bones, tissues, and brain of the body in a technique called a Positron Emission Topography scan (PET scan). An example of an ^{18}F -PET scan is displayed (figure A1). In this scan, ^{18}F is being used to observe the differences in brain activity between a sober and an intoxicated brain. With the half-life of 110 minutes, there is little chance of radiation damage to the patient because the amount of radioactive fluorine will decrease by a factor of about 10^4 within one day. Selected applications and selected uses of stable and (or) radioactive isotopes in our everyday lives for each of the chemical elements will be featured.

Appendix A

Example Element Page—Fluorine

fluorine

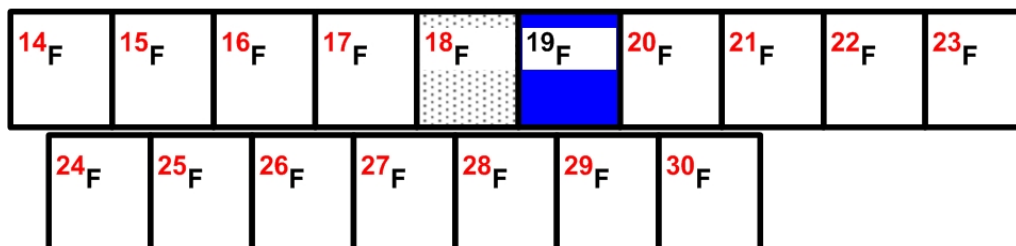


Stable isotope	Atomic mass*	Mole fraction
^{19}F	18.998 403 22	1.0000

* Atomic mass given in unified atomic mass units, u.

Half-life of radioactive isotope

Less than 1 hour
 Between 1 hour and 1 year
 Greater than 1 year



Selected Applications of Stable and/or Radioactive Isotopes of Fluorine

Fluorine Isotopes in Medicine

- 1) ^{18}F is a radioactive fluorine isotope that is used in a ^{18}F -FDG compound (^{18}F -labeled, fluoro-deoxy glucose) for imaging the organs, bones, tissues and brain of the body with a technique called a Positron Emission Topography scan (PET).
 - a. ^{18}F emits positrons (positive electrons) that collect in tissue and interact with regular negative electrons when injected into the body. The positrons and electrons annihilate each other, producing two gamma particles that are emitted in opposite directions and this causes the release of X-ray-like radiation. The radiation is detected on a PET camera, which generates a picture of the body part being examined.

- b. Because ^{18}F has a short half-life of about 110 minutes, there is little chance of radiation damage to the patient.

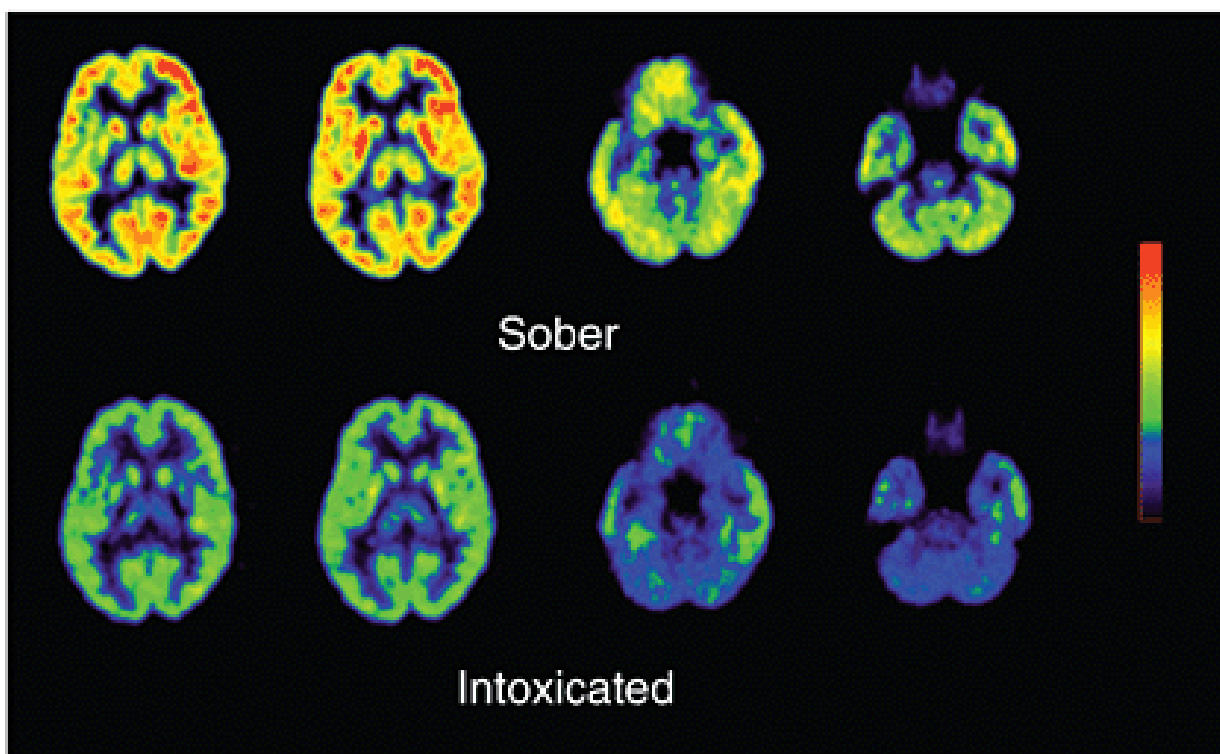


Figure A1: An ^{18}F -FDG PET scan is used to observe the differences in brain activity between a sober and an intoxicated brain. (Image Source: National Institute on Alcohol Abuse and Alcoholism (NIAAA)).

References

Fluorine

Fluorine Isotopes in Medicine

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Element Page Last Modified: May 4, 2012