IUPAC Periodic Table of the Isotopes

- Element has two or more stable isotopes. Atomic weight and isotopic abundances of element vary in naturally occurring materials. The lower and upper bounds of atomic weight have been assessed by IUPAC and are presented as the standard atomic weight within square brackets, [ ].
- Element has two or more stable isotopes and the standard atomic weight is not a constant of nature. The lower and upper bounds of the standard atomic weight have not been evaluated by IUPAC yet.
- Element has one stable isotope and its standard atomic weight is a constant of nature.
- Element has no stable isotopes. Thus, no standard atomic weight exists.

-Cd
- Stable isotope mass numbers (number of protons + neutrons)
- Isotopic abundances (mass fractions of stable isotopes)
- Uncertainty in last digit

Boron (B)
Carbon (C)
Nitrogen (N)
Oxygen (O)
Fluorine (F)
Neon (Ne)
Aluminum (Al)
Silicon (Si)
Phosphorus (P)
Sulfur (S)
Chlorine (Cl)
Argon (Ar)
Potassium (K)
Calcium (Ca)
Scandium (Sc)
Titanium (Ti)
Vanadium (V)
Chromium (Cr)
Manganese (Mn)
Iron (Fe)
Co (Co)
Cobalt (Co)
Nickel (Ni)
Copper (Cu)
Zinc (Zn)
Gallium (Ga)
Ge (Ge)
Arsenic (As)
Se (Se)
Br (Br)
Krypton (Kr)
Rubidium (Rb)
Strontium (Sr)
Cesium (Cs)
Barium (Ba)
Lanthanum (La)
Cerium (Ce)
Praseodymium (Pr)
Neodymium (Nd)
Promethium (Pm)
Samarium (Sm)
Europium (Eu)
Gadolinium (Gd)
Terbium (Tb)
Dysprosium (Dy)
Holmium (Ho)
Erbium (Er)
Thulium (Tm)
Ytterbium (Yb)
Lutetium (Lu)

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The number of protons in each atom (i.e., the atomic number, symbol Z) determines the chemical properties of that atom. All atoms of a given element have the same number of protons (e.g., carbon atoms have 6 protons, while all uranium atoms have 92 protons). The number of neutrons (symbol N) in an atom of a given element may vary. Atoms of a given element which have different numbers of neutrons are called isotopes.

The total number of protons and neutrons (Z + N) in a given isotope is called the mass number.

Isotopes may be stable or unstable. The unstable isotopes are also called radio-isotopes and will decay over time into another isotope of the same or a different element. The stable isotopes of each element and their mass numbers appear on each element cell of the Table in a pie chart, where each mass number is indicated around the outside of the pie. The relative amount of each stable isotope in a given element (called the isotopic abundance) is approximately indicated by the size of the slice of the pie assigned to each mass number.

Some elements have no stable isotopes and therefore no mass numbers; these elements are indicated in white on the Table. Other elements have only one stable isotope and a single mass number and they are indicated in blue. The remaining elements have more than one stable isotope and mass number; these elements are indicated in either yellow or pink.

The atomic weight of an element is calculated from the sum of the products of the atomic mass and the isotopic abundance of each stable isotope of that element. The Standard Atomic Weight is the recommended value that can apply to all samples of a given element. Consider the simplified calculation for the case of carbon (all of whose isotopes have 6 protons since its atomic number is 6). Carbon has 2 stable isotopes of mass number 12 (abbreviated as 12C; corresponding to an atom with 6 protons and 6 neutrons) and mass number 13 (13C; corresponding to an atom with 6 protons and 7 neutrons). The atomic mass of each isotope can be approximated by its mass number, 1C = 12 and 13C = 13. Natural carbon is a mixture of 12C and 13C atoms with appropriate isotopic abundances of 98.9% and 1.1%, respectively. The approximate atomic weight for this sample of carbon would be 12 x 0.989 + 13 x 0.011 = 12.01.

For any element that has two or more stable isotopes, there is always the possibility that the relative amounts of stable isotopes may vary in various samples of that element found in nature. Using the above example, let us assume that another sample of carbon is made up of 98% 12C and 2% 13C. This sample of carbon has an approximate atomic weight of 12 x 0.98 + 13 x 0.02 = 12.02. Thus, natural isotopic variation for an element can have an effect on the element’s atomic weight value. For 10 such elements, the Standard Atomic Weight assigned by IUPAC is given as upper and lower bounds (called an interval) written in brackets (e.g., for chlorine, it is [35.446; 35.457]). These elements are indicated in pink on the Table. Those elements for which no such assessments have been made or completed yet are indicated in yellow and their Standard Atomic Weights are given with an uncertainty in parentheses (e.g., for mercury, 200.59(2) is a contracted notation of 200.59 +/- 0.02).

References

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