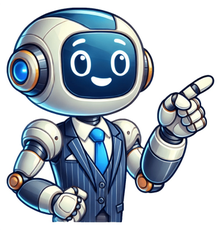
I'm not a robot



Some amount of mass is associated with every particle of matter, no matter if it’s small or large in size. Everything around us is composed of atoms, and the mass of these atoms is known as atomic mass. This is generally expressed in terms of atomic mass unit (amu) and can be measured almost accurately with the help of spectrometry. The carbon atomic mass is also measured relative to the mass of a standardized atom by experimental means. Hydrogen is the first element in the periodic table and has an atomic number of 1, so its relative atomic mass is considered to be 1. The carbon atomic mass is described as the 1/12th of the mass of a single carbon-12 atom, and it equals to 1.66*10-24 grams approximately. Isotope abundance (relative to each other that exists on Earth) is taken into account to calculate the average atomic mass. Suppose carbon-12 is the most abundantly found isotope of carbon on Earth. Its relative abundance is almost equal to 98.89%. 13.003 and 14.003 are the atomic masses of carbon 13 and carbon 14, respectively. From these isotopes, we calculate the average atomic mass of carbon as a whole which comes out to be equal to 12.011 amu. Atomic Mass As discussed earlier, the atomic mass of sodium and other metals is defined as being equal to 1/12th of carbon-12 atom mass when it is in its ground state. The total average atomic mass of an atom is equal to the sum of the individual masses of its neutrons and protons. When the atomic weight of an atom is divided by the unified atomic weight in Daltons, it is a dimensionless number. It is known as the relative isotopic mass. The atomic masses of different elements are different. The variation is so much so that the atomic mass of the lightest element hydrogen atom, is 1.008, while that of the heaviest atom is 250 amu. The average mass of a molecule can be calculated by combining the average atomic mass of each individual atom it is made up of. Atomic Mass of Elements Below is a table showing the average atomic mass of different elements: ElementAtomic MassChlorine35.453 uIron55.845 uCopper63.546 uSulphur32.065 uNitrogen14.0067 uPotassium39.0983 uPhosphorus30.973762 uCalcium40.078 uSodium22.989769 u Average Mass Atoms with the same atomic numbers but different mass numbers are called isotopes. The different mass numbers are because they contain different numbers of neutrons. The actual masses of the individual atoms are never whole numbers, although the mass numbers are. This is how lithium has an atomic number of 6.942 Da. The isotopes of different atoms in the periodic table are arranged in order according to their atomic masses and abundance in nature. The isotopes that are more abundant in nature are given more weight. An element’s average mass is defined as: m(E)=∑ m(In)×p(In), n=1 Where ∑ indicates an n-times summation of all the isotopes In of element E, and p(I) indicates the isotope I’s relative abundance. Relative Mass Traditionally when atomic masses were used in Chemistry, usage of any units was avoided. Even nowadays, a chemist will say that carbon-12 has a mass exactly equal to 12. However, as mass is not a dimensional entity, it is understood that it requires a unit. The omission of the units used to be rationalized because of the concept of relative mass, which is a ratio of two masses instead of a single mass. For relative mass, the ratio of the mass is used rather than using a unit. This is constant mu is used analogously to the unit that is being avoided. Therefore, we can define the relative atomic mass of sodium and other metals as: Ar = m/mu Now, this is a dimensionless quantity. As the representation of relative mass is against the standards of modern metrology, its use is generally discouraged. Integer Mass Because protons and neutrons have almost the same mass, and the mass of an electron is very low compared to both of them, the mass of most molecules, if measured in daltons, is close to an integer value. Therefore, very commonly, only the integer mass of molecules is indicated. However, it is only meaningful if units of dalton are used. Accurate Mass The mass of a molecule can be determined with an accuracy higher than the integer mass by many mass spectrometers. Therefore, this measurement is called the molecule’s accurate mass. The atomic masses and integer masses of isotopes are not identical due to a mass defect that results from the binding energy in the nucleus. What is Molecular Mass? A compound’s molecular mass can be calculated by the sum of the atomic masses of the individual atoms that it is made up of. Molecular mass is sometimes also known as molecular weight. For instance, there is one oxygen atom and two hydrogen atoms in a molecule of water. Its mass equals the average mass of an atom of hydrogen multiplied by two and added by the oxygen atomic mass. An element’s molecular mass is determined by the atomic masses of the atoms constituting it. Following are the methods by which the molecular mass of different molecules can be measured: Mass spectrometry: The masses of small molecules can easily be measured by this method and reported as monoisotopic mass. Hydrodynamic method: Mark-Houwink relations are used to determine the molecular weights in this method. Calibration is needed in this method; therefore, it is also defined as a method to determine the relative molecular weight. Static Light Scattering: The measurement of the amount of light scattered by the Zimm method is used to determine the molecular weight. Key Points Atoms are the basic building blocks of all matter. Substances that are made up of a single atom are called elements. Every atom has a nucleus at the center, which is made up of neutrons and protons. Electrons revolve around the nucleus in orbits. The relative mass of protons is one, and those are positively charged. The relative mass of neutrons is also one, but those are neutrally charged. The mass of an electron is so small that it is almost negligible. It is negatively charged. Conclusion Hope that the article helps to clear your doubts about the concept of carbon atomic mass, the atomic mass of calcium, the atomic mass of sodium and the atomic mass of other elements. As it is an important concept in both Physics and Chemistry, having a fundamental understanding of its concepts is important. Frequently Asked Questions 1. Calculate the molar mass of sulphur trioxide. The definition of molar mass suggests that it is equal to the molecular mass in grams. So we should first calculate its molecular mass. Solution: A compound’s molecular mass can be defined as the sum of the atomic masses of its constituent atoms. Following is the molecular mass of sulphur trioxide with four atoms. 1S = 1 x 32.06 amu = 32.06 amu 3O = 3 x 16.00 amu = 48.00amu Therefore, 80.07 amu is the molecular mass of sulphur trioxide. 2. Define the atomic mass of calcium in simple terms. In simple terms, the atomic mass of calcium can be defined as the mass of a single atom of a chemical element. It is a combination of the masses of three subatomic particles, namely neutrons, protons, and electrons. 3. Mention the importance of oxygen atomic mass. Oxygen atomic mass is a vital concept in Chemistry as it establishes a connection between mass and moles that is measured in the laboratory. Almost everything you study in Chemistry depends on the properties of an atom, including its mass. 4. Define atomic mass number. This number can be defined as the combination of a total number of protons and a total number of neutrons in an atom. It is denoted by the symbol A. 5. How can molecular mass be determined? Molecular mass is an entity that is equal to the number of nuclear masses of a molecule. It provides the weight of a molecule relative to that of the most abundant isotope of carbon-carbon 12, which is considered to have a weight of 12. 6. Is gold a heavy element? Even though gold is soft and malleable, it is considered to be a heavy metal as each of its constituent atoms is heavy. 7. Mention the fundamental unit of mass. The concept of mass is used to calculate the weight of an object. Suppose you can determine the mass of your body when you step on a scale. Kilograms and grams are the most important mass units in the metric system. La masa atómica es un concepto fundamental en la química y la física, y es especialmente importante en la tabla periódica. En este artículo, exploraremos qué es la masa atómica, cómo se mide y por qué es tan importante en la tabla periódica.¿Qué verás en este artículo? La masa atómica es la masa de un átomo de un elemento en particular. Cada átomo tiene una masa atómica única que se expresa en unidades de masa atómica (uma) o kilogramos (kg). La masa atómica se calcula sumando el número de protones y neutrones en el núcleo de un átomo.¿Cómo se mide la masa atómica?La masa atómica se mide utilizando la unidad de masa atómica (uma). Una unidad de masa atómica es igual a la masa de un átomo de carbono-12, que se toma como referencia. La masa atómica de un átomo se expresa en términos de una fracción de la masa de un átomo de carbono-12.¿Por qué es importante la masa atómica en la tabla periódica?La tabla periódica se organiza en función de la masa atómica de los elementos. Los elementos se ordenan en filas y columnas en función de sus propiedades químicas y físicas. La masa atómica es un factor importante en la determinación de estas propiedades.¿Cómo se organizan los elementos en la tabla periódica?Los elementos se organizan en la tabla periódica en función de su número atómico y su configuración electrónica. Los elementos se agrupan en filas horizontales llamadas períodos y en columnas verticales llamadas grupos. Los elementos en el mismo grupo tienen propiedades químicas y físicas similares.¿Cómo afecta la masa atómica a las propiedades de los elementos?La masa atómica afecta a varias propiedades de los elementos, incluyendo su densidad, punto de fusión y punto de ebullición. Los elementos con una masa atómica más alta tienden a ser más densos y tener puntos de fusión y ebullición más altos que los elementos con una masa atómica más baja.¿Cómo se relacionan la masa atómica y el tamaño del átomo?A medida que aumenta la masa atómica de un elemento, también aumenta el tamaño del átomo. Esto se debe a que el número de protones y electrones aumenta, lo que aumenta la atracción entre el núcleo y los electrones y hace que el átomo se expanda.ConclusiónLa masa atómica es una propiedad fundamental de los átomos y es clave en la organización de la tabla periódica. La masa atómica afecta a muchas propiedades físicas y químicas de los elementos y es un factor importante a considerar en la investigación y aplicación de los elementos en diferentes campos.Preguntas frecuentes¿Por qué algunos elementos tienen una masa atómica no entera?La masa atómica de algunos elementos no es un número entero porque es una media ponderada de las masas isotópicas de los átomos del elemento. Los isótopos son átomos del mismo elemento con diferentes números de neutrones, lo que afecta la masa atómica.¿Qué es la masa molar?La masa molar es la masa de un mol de una sustancia, que se expresa en gramos. La masa molar se calcula sumando las masas atómicas de todos los átomos en una molécula o fórmula química.¿Cómo se determina la masa atómica de un elemento?La masa atómica de un elemento se determina mediante una técnica llamada espectrometría de masas. En esta técnica, se ioniza un átomo o molécula y se mide su relación masa-carga para determinar su masa atómica.¿Por qué la masa atómica del hidrógeno es 1.008?La masa atómica del hidrógeno es 1.008 porque el hidrógeno tiene dos isótopos, el hidrógeno-1 y el hidrógeno-2 (deuterio). La masa atómica del hidrógeno es una media ponderada de las masas isotópicas de estos dos isótopos.¿Cómo se relaciona la masa atómica con la energía de un átomo?La masa atómica afecta a la energía de un átomo a través de la ecuación de Einstein, E = mc2. Esta ecuación establece que la energía (E) de un objeto es igual a su masa (m) multiplicada por la velocidad de la luz (c) al cuadrado. Por lo tanto, la masa atómica afecta a la energía de un átomo y su capacidad para interactuar con otros átomos y moléculas. views updated May 14 2018Every known item of matter in the universe has some amount of mass, even if it is very small. But what about something so insignificant in mass that comparing it to a gram is like comparing a millimeter to the distance between Earth and the nearest galaxy? Obviously, special units are needed for such measurements; then again, one might ask why it is necessary to weigh atoms at all. One answer is that everything is made of atoms. More specifically, the work of a chemist requires the use of accurate atomic proportions in forming the molecules that make up a compound. The measurement of atomic mass was thus a historic challenge that had to be overcome, and the story of the ways that scientists met this challenge is an intriguing one.HOW IT WORKSWhy Mass and Not Weight?Some textbooks and other sources use the term atomic weight instead of atomic mass. The first of these is not as accurate as the second, which explains why atomic mass was chosen as the subject of this essay. Indeed, the use of "atomic weight" today merely reflects the fact that scientists in the past used that expression and spoke of "weighing" atoms. Though "weigh" is used as a verb in this essay, this is only because it is less cumbersome than "measure the mass of." (In addition, "atomic weight" may be mentioned when discussing studies by scientists of the nineteenth century, who applied that term rather than atomic mass.)One might ask why such pains have been taken to make the distinction. Mass is, after all, basically the same as weight, is it not? In fact it is not, though people are accustomed to thinking in those terms since most weight scales provide measurements in both pounds and kilograms. However, the pound is a unit of weight in the English system, whereas a kilogram is a unit of mass in the metric and SI systems. Though the two are relatively convertible on Earth, they are actually quite different. (Of course it would make no more sense to measure atoms in pounds or kilograms than to measure the width of a hair in light-years; but pounds and kilograms are the most familiar units of weight and mass respectively.)Weight is a measure of force affected by Earth's gravitational pull. Therefore a person's weight varies according to gravity, and would be different if measured on the Moon, whereas mass is the same throughout the universe. Its invariability makes mass preferable to weight as a parameter of scientific measure.Putting an Atom's Size and Mass in ContextMass does not necessarily relate to size, though there is enough of a loose correlation that more often than not, we can say that an item of very small size will have very small mass. And atoms are very, very small—so much so that, until the early twentieth century, chemists and physicists had no accurate means of isolating them to determine their mass.The diameter of an atom is about 10−8 cm. This is equal to about 0.000000003937 in—or to put it another way, an inch is about as long as 250 million atoms lined up side by side. Obviously, special units are required for describing the size of atoms. Usually, measurements are provided in terms of the angstrom, equal to 10−10 m. (In other words, there are 10 million angstroms in a millimeter.)Measuring the spatial dimensions of an atom, however, is not nearly as important for chemists' laboratory work as measuring its mass. The mass of an atom is almost inconceivably small. It takes about 5.0 · 1023 carbon atoms to equal just one gram in mass. At first, 1023 does not seem like such a huge number, until one considers that 106 is already a million, meaning that 1023 is a million times a million times a million times 100,000. If 5.0 · 1023 angstrom lengths—angstroms, not meters or even millimeters—were laid end to end, they would stretch from Earth to the Sun and back 107,765 times!Atomic Mass UnitsIt is obvious, then, that an entirely different unit is needed for measuring the mass of an atom, and for this purpose, chemists and other scientists use an atom mass unit (abbreviated amu), which is equal to 1.66 · 10−24 g.Though the abbreviation amu is used in this book, atomic mass units are sometimes designated simply by a u. On the other hand, they may be presented as numbers without any unit of measure—as for instance on the periodic table of elements.Within the context of biochemistry and microbiology, often the term dalton (abbreviated Da or D) is used. This is useful for describing the mass of large organic molecules, typically rendered in kilodaltons (kDa). The Latin prefix kilo-indicates 1,000 of something, and "kilodalton" is much less of a tongue-twister than "kilo-amu". The term "dalton" honors English chemist John Dalton (1766-1844), who, as we shall see, introduced the concept of the atom to science.AVERAGE ATOMIC MASS.Since 1960, when its value was standardized, the atomic mass unit has been officially known as the "unified atomic mass unit." The addition of the word "unified" reflects the fact that atoms are not weighed individually—a labor that would be problematic at the very least. In any case, to do so would be to reinvent the wheel, as it were, because average atomic mass figures have been established for each element.Average atomic mass figures range from 1.008 amu for hydrogen, the first element listed on the periodic table of elements, to over 250 amu for elements of very high atomic number. Figures for average atomic mass can be used to determine the average mass of a molecule as well, since a molecule is just a group of atoms joined in a structure. The mass of a molecule can be determined simply by adding together average atomic mass figures for each atom the molecule contains. A water molecule, for instance, consists of two hydrogen atoms and one oxygen atom; therefore, its mass is equal to the average atomic mass of hydrogen multiplied by two, and added to the average atomic mass of oxygen.Avogadro's Number and the MoleAtomic mass units and average atomic mass are not the only components necessary for obtaining accurate mass figures where atoms are concerned. Obviously, as suggested several times already, it would be fruitless to determine the mass of individual atoms or molecules. Nor would it do to measure the mass of a few hundred, or even a few million, of these particles. After all, as we have seen, it takes about 500,000 trillion million carbon atoms to equal just one gram—and a gram, after all, is still rather small in mass compared to most objects encountered in daily life. (There are 1,000 g in a kilogram, and a pound is equal to about 454 g.)In addition, scientists need some means for comparing atoms or molecules of different substances in such a way that they know they are analyzing equal numbers of particles. This cannot be done in terms of mass, because the number of atoms in each sample varies: a gram of hydrogen, for instance, contains about 12 times as many atoms as a gram of carbon, which has an average atomic mass of 12.01 amu. What is needed, instead, is a way to designate a certain number of atoms or molecules, such that accurate comparisons are possible.In order to do this, chemists make use of a figure known as Avogadro's number. Named after Italian physicist Amedeo Avogadro (1776-1856), it is equal to 6.022137 × 1023. Avogadro's number, which is 6,022,137 followed by 17 zeroes, designates the quantity of molecules in a mole (abbreviated mol). The mole, a fundamental SI unit for "amount of substance," is defined precisely as the number of carbon atoms in 12.01 g of carbon. It is here that the value of Avogadro's number becomes clear: as noted, carbon has an average atomic mass of 12.01 amu, and multiplication of the average atomic mass by Avogadro's number yields a figure in grams equal to the value of the average atomic mass in atomic mass units.REAL-LIFE APPLICATIONSEarly Ideas of Atomic MassDalton was not the first to put forth the idea of the atom: that concept, originated by the ancient Greeks, had been around for more than 2,000 years. However, atomic theory had never taken hold in the world of science—or, at least, what passed for science prior to the seventeenth century revolution in thinking brought about by Galileo Galilei (1564-1642) and others.Influenced by several distinguished predecessors, Dalton in 1803 formulated the theory that nature is formed of tiny particles, an idea he presented in A New System of Chemical Philosophy (1808). Dalton was the first to treat atoms as fully physical constructs; by contrast, ancient proponents of atomism conceived these fundamental particles in ideal or spiritual terms. Dalton described atoms as hard, solid, indivisible particles with no inner spaces—a definition that did not endure, as later scientific inquiry revealed the complexities of the atom. Yet he was correct in identifying atoms as having weight—or, as scientists say today, mass.THE FIRST TABLE OF ATOMIC WEIGHTS.The question was, how could anyone determine the weight of something as small as an atom? A year after the publication of Dalton's book, a discovery by French chemist and physicist Joseph Gay-Lussac (1778-1850) and German naturalist Alexander von Humboldt (1769-1859) offered a clue. Humboldt and Gay-Lussac—famous for his gas law associating pressure and temperature—found that gases combine to form compounds in simple proportions by volume.For instance, as Humboldt and Gay-Lussac discovered, water is composed of only two elements: hydrogen and oxygen, and these two combine in a whole-number ratio of 8:1. By separating water into its components, they found that for every part of oxygen, there were eight parts of hydrogen. Today we know that water molecules are formed by two hydrogen atoms, with an average atomic mass of 1.008 amu each, and one oxygen atom. The ratio between the average atomic mass of oxygen (16.00 amu) and that of the two hydrogen atoms is indeed very nearly 8:1.In the early nineteenth century, however, chemists had no concept of molecular structure, or any knowledge of the atomic masses of elements. They could only go on guesswork: hence Dalton, in preparing the world's first "Table of Atomic Weights," had to make some assumptions based on Humboldt's and Gay-Lussac's findings. Presumably, Dalton reasoned, only one atom of hydrogen combines with one atom of oxygen to form a "water atom." He assigned to hydrogen a weight of 1, and according to this, calculated the weight of oxygen as 8.AVOGADRO AND BERZELIUS IMPROVE ON DALTON'S WORK.The implications of Gay-Lussac's discovery that substances combined in whole-number ratios were astounding. (Gay-Lussac, who studied gases for much of his career, is usually given more credit than Humboldt, an explorer and botanist who had his hand in many things.) On the one hand, the more scientists learned about nature, the more complex it seemed; yet here was something amazingly simple. Instead of combining in proportions of, say, 8.3907 to 1.4723, oxygen and hydrogen molecules formed a nice, clean, ratio of 8 to 1. This served to illustrate the fact that, as Dalton had stated, the fundamental particles of matter must be incredibly tiny; otherwise, it would be impossible for every possible quantity of hydrogen and oxygen in water to have the same ratio.Intrigued by the work of Gay-Lussac, Avogadro in 1811 proposed that equal volumes of gases have the same number of particles if measured at the same temperature and pressure. He also went on to address a problem raised by Dalton's work. If atoms were indivisible, as Dalton had indicated, how could oxygen exist both as its own atom and also as part of a water "atom"? Water, as Avogadro correctly hypothesized, is not composed of atoms but of molecules, which are themselves formed by the joining of two hydrogen atoms with one oxygen atom.Avogadro's molecular theory opened the way to the clarification of atomic mass and the development of the mole, which, as we have seen, makes it possible to determine mass for large quantities of molecules. However, his ideas did not immediately gain acceptance. Only in 1860, four years after Avogadro's death, did Italian chemist Stanislao Cannizzaro (1826-1910) resurrect the concept of the molecule as a way of addressing disagreements among scientists regarding the determination of atomic mass.In the meantime, Swedish chemist Jons Berzelius (1779-1848) had adopted Dalton's method of comparing all "atomic weights" to that of hydrogen. In 1828, Berzelius published a table of atomic weights, listing 54 elements along with their weight relative to that of hydrogen. Thus carbon, in Berzelius's system, had a weight of 12. Unlike Dalton's figures, Berzelius's are very close to those used by scientists today. By the time Russian chemist Dmitri Ivanovitch Mendeleev (1834-1907) created his periodic table in 1869, there were 63 known elements. That first table retained the system of measuring atomic mass in comparison to hydrogen.The Discovery of Subatomic StructuresUntil scientists began to discover the existence of subatomic structures, measurements of atomic mass could not really progress. Then in 1897, English physicist J. J. Thomson (1856-1940) identified the electron. A particle possessing negative charge, the electron contributes little to an atom's mass, but it pointed the way to the existence of other particles within an atom. First of all, there had to be a positive charge to offset that of the electron, and secondly, the item or items providing this positive charge had to account for the majority of the atom's mass.Early in the twentieth century, Thomson's student Ernest Rutherford (1871-1937) discovered that the atom has a nucleus, a center around which electrons move, and that the nucleus contains positively charged particles called protons. Protons have a mass 1,836 times as great as that of an electron, and thus seemed to account for the total atomic mass. Later, however, Rutherford and English chemist Frederick Soddy (1877-1956) discovered that when an atom emitted certain types of particles, its atomic mass changed.ISOTOPES AND ATOMIC MASS.Rutherford and Soddy named these atoms of differing mass isotopes, though at that point—because the neutron had yet to be discovered—they did not know exactly what had caused the change in mass. Certain types of isotopes, Soddy and Rutherford concluded, had a tendency to decay, moving (sometimes over a great period of time) toward stabilization. Such isotopes were radioactive.Soddy concluded that atomic mass, as measured by Berzelius, was actually an average of the mass figures for all isotopes within that element. This explained a problem with Mendeleev's periodic table, in which there seemed to be irregularities in the increase of atomic mass from element to element. The answer to these variations in mass, it turned out, related to the number of isotopes associated with a given element: the greater the number of isotopes, the more these affected the overall measure of the element's mass.A NEW DEFINITION OF ATOMIC NUMBER.Up to this point, the term "atomic number" had a different, much less precise, meaning than it does today. As we have seen, the early twentieth century periodic table listed elements in order of their atomic mass in relation to hydrogen, and thus atomic number referred simply to an element's position in this ordering. Then, just a few years after Rutherford and Soddy discovered isotopes, Welsh physicist Henry Moseley (1887-1915) determined that every element has a unique number of protons in its nucleus.Today, the number of protons in the nucleus, rather than the mass of the atom, determines the atomic number of an element. Carbon, for instance, has an atomic number of 6, not because there are five elements lighter—though this is also true—but because it has six protons in its nucleus. The ordering by atomic number happens to correspond to the ordering by atomic mass, but atomic number provides a much more precise means of distinguishing elements. For one thing, atomic number is always a whole integer—1 for hydrogen, for instance, or 17 for chlorine, or 92 for uranium. Figures for mass, on the other hand, are almost always rendered with decimal fractions (for example, 1.008 for hydrogen).NEUTRONS COMPLETE THE PICTURE.As with many other discoveries along the way to uncovering the structure of the atom, Moseley's identification of atomic number with the proton raised still more questions. In particular, if the unique number of protons identified an element, what was it that made isotopes of the same element different from one another? Hydrogen, as it turned out, indeed had a mass very nearly equal to that of one proton—thus justifying its designation as the basic unit of atomic mass. Were it not for the isotope known as deuterium, which has a mass nearly twice as great as that of hydrogen, the element would have an atomic mass of exactly 1 amu.A discovery by English physicist James Chadwick (1891-1974) in 1932 finally explained what made an isotope an isotope. It was Chadwick who identified the neutron, a particle with no electric charge, which resides in the nucleus alongside the protons. In deuterium, which has one proton, one neutron, and one electron, the electron accounts for only 0.0272% of the total mass—a negligible figure. The proton, on the other hand, makes up 49.9392% of the mass. Until the discovery of the neutron, there had been no explanation of the other 50.0336% of the mass in an atom with just one proton and one electron.Average Atomic Mass TodayThanks to Chadwick's discovery of the neutron, it became clear why deuterium weighs almost twice as much as ordinary hydrogen. This in turn is the reason why a large sample of hydrogen, containing as it does a few molecules of deuterium here and there, does not have the same average atomic mass as a proton. Today scientists know that there are literally thousands of isotopes—many of them stable, but many more of them unstable or radioactive—for the 100-plus elements on the periodic table. Each isotope, of course, has a slightly different atomic mass. This realization has led to clarification of atomic mass figures.One might ask how figures of atomic mass are determined. In the past, as we have seen, it was largely a matter of guesswork, but today chemists and physicist use a highly sophisticated instrument called a mass spectrometer. First, atoms are vaporized, then changed to positively charged ions, or cations, by "knocking off" electrons. The cations are then passed through a magnetic field, and this causes them to be deflected by specific amounts, depending on the size of the charge and its atomic mass. The particles eventually wind up on a deflector plate, where the amount of deflection can be measured and compared with the charge. Since 1 amu has been calculated to equal approximately 931.494 MeV, or mega electron-volts, very accurate figures can be determined.CALIBRATION OF THE ATOMIC MASS UNIT.When 1 is divided by Avogadro's number, the result is 1.66 · 10−24—the value, in grams, of 1 amu. However, in accordance with a 1960 agreement among members of the international scientific community, measurements of atomic mass take as their reference point the mass of carbon-12. Not only is the carbon-12 isotope found in all living things, but hydrogen is a problematic standard because it bonds so readily with other elements. According to the 1960 agreement, 1 amu is officially 1/12 the mass of a carbon-12 atom, whose exact value (retested in 1998), is 1.6653873 · 10−24 g.Carbon-12, sometimes represented as contains six protons and six neutrons. (As explained in the essay on Isotopes, where an isotope is indicated, the number to the upper left of the chemical symbol indicates the total number of protons and neutrons. Sometimes this is the only number shown; but if a number is included on the lower left, this indicates only the number of protons, which remains the same for each element.) The value of 1 amu thus obtained is, in effect, an average of the mass for a proton and neutron—a usable figure, given the fact that a neutron weighs only 0.163% more than a proton.Of all the carbon found in nature (as opposed to radioactive isotopes created in laboratories), 98.89% of it is carbon-12. The remainder is mostly carbon-13, with traces of carbon-14, an unstable isotope produced in nature. By definition, carbon-12 has an atomic mass of exactly 12 amu; that of carbon-13 (about 1.11% of all carbon) is 13 amu. Thus the atomic mass of carbon, listed on the periodic table as 12.01 amu, is obtained by taking 98.89% of the mass of carbon-12, combined with 1.11% of the mass of carbon-13.ATOMIC MASS UNITS AND THE PERIODIC TABLE.The periodic table as it is used today includes figures, in atomic mass units, for the average mass of each atom. As it turns out, Berzelius was not so far off in his use of hydrogen as a standard, since its mass is almost exactly 1 amu—but not quite, because (as noted above) deuterium increases the average mass somewhat. Figures increase from there along the periodic table, though not by a regular pattern. Sometimes the increase from one element to the next is by just over 1 amu, and in other cases, the increase is by more than 3 amu. This only serves to prove that atomic number, rather than atomic mass, is a more straightforward means of ordering the elements.Mass figures for many elements that tend to appear in the form of radioactive isotopes are usually shown in parentheses. This is particularly true for elements with atomic numbers above 92, because samples of these elements do not stay around long enough to be measured. Some have a half-life—the period in which half the isotopes decay to a stable form—of just a few minutes, and for others, the half-life is but a fraction of a second. Therefore, atomic mass figures represent the mass of the longest-lived isotope.Uses of Atomic Mass in ChemistryMOLAR MASS.Just as the value of atomic mass units has been calibrated to the mass of carbon-12, the mole is no longer officially defined in terms of Avogadro's number, though in general its value has not changed. By international scientific agreement, the mole equals the number of carbon atoms in 12.01 g of carbon. Note that, as stated earlier, carbon has an average atomic mass of 12.01 amu.This is no coincidence, of course: multiplication of the average atomic mass by Avogadro's number yields a figure in grams equal to the value of the average atomic mass in amu. A mole of helium, with an average atomic mass of 4.003, is 4.003 g. Iron, on the other hand, has an average atomic mass of 55.85, so a mole of iron is 55.85 g. These figures represent the molar mass—the mass of 1 mole—for each of the elements mentioned.THE NEED FOR EXACT PROPORTIONS.When chemists discover new substances in nature or create new ones in the laboratory, the first thing they need to determine is the chemical formula—in other words, the exact quantities and proportions of elements in each molecule. By chemical means, they separate the compound into its constituent elements, then determine how much of each element is present.Since they are using samples in relatively large quantities, molar mass figures for each element make it possible to determine the chemical composition. To use a very simple example, suppose a quantity of water is separated, and the result is 2.016 g of hydrogen and 16 g of oxygen. The latter is the molar mass of oxygen, and the former is the molar mass of hydrogen multiplied by two. Thus we know that there are two moles of hydrogen and one mole of oxygen, which combine to make one mole of water.Of course the calculations used by chemists working in the research laboratories of universities, government institutions, and corporations are much, much more complex than the example we have given. In any case, it is critical that a chemist be exact in making these determinations, so as to know the amount of reactants needed to produce a given amount of product, or the amount of product that can be produced from a given amount of reactant.When a company produces millions or billions of a single item in a given year, a savings of very small quantities in materials—thanks to proper chemical measurement—can result in a savings of billions of dollars on the bottom line. Proper chemical measurement can also save lives. Again, to use a very simple example, if a mole of compounds weighs 44.01 g and is found to contain two moles of oxygen and one of carbon, then it is merely carbon dioxide—a compound essential to plant life. But if it weighs 28.01 g and has one mole of oxygen with one mole of carbon, it is poisonous carbon monoxide.WHERE TO LEARN MORE"Atomic Weight" (Web site). < (May 23, 2001)."An Experiment with 'Atomic Mass'" (Web site). < (May 23, 2001).Knapp, Brian J. and David Woodroffe. The Periodic Table. Danbury, CT: Grolier Educational, 1998.Oxlade, Chris. Elements and Compounds. Chicago: Heinemann Library, 2001."Periodic Table: Atomic Mass." ChemicalElements.com (Web site). < (May 23, 2001)."Relative Atomic Mass" (Web site). < (May 23, 2001)."What Are Atomic Number and Atomic Weight?" (Website). < (May 23, 2001).KEY TERMSATOM:The smallest particle of an element that retains all the chemical and physical properties of that element.ATOMIC MASS UNIT:An SI unit (abbreviated amu), equal to 1.66 · 10−24 g, for measuring the mass of atoms.ATOMIC NUMBER:The number of protons in the nucleus of an atom. Since this number is different for each element, elements are listed on the periodic table of elements in order of atomic number.ATOMIC WEIGHT:An old term for atomic mass. Since weight varies depending on gravitational field, whereas mass is the same throughout the universe, scientists typically use the term "atomic mass" instead.AVERAGE ATOMIC MASS:A figure used by chemists to specify the mass—in atomic mass units—of the average atom in a large sample.AVOGADRO'S NUMBER:A figure, named after Italian physicist Amedeo Avogadro (1776-1856), equal to 6.022137 × 1023. Avogadro's number indicates the number of atoms or molecules in a mole.COMPOUND:A substance made up of atoms of more than one element. These atoms are usually joined in molecules.DALTON:An alternate term for atomic mass units, used in biochemistry and microbiology for describing the mass of large organic molecules. The dalton (abbreviated Da or D) is named after English chemist John Dalton (1766-1844), who introduced the concept of the atom to science.ELEMENT:A substance made up of only one kind of atom, which cannot be chemically broken into other substances.HALF-LIFE:The length of time it takes a substance to diminish to one-half its initialamount.ION:An atom or group of atoms that has lost or gained one or more electrons, and thus has a net electric charge.ISOTOPES:Atoms of the same element (that is, they have the same number of protons) that differ in terms of mass. Isotopes may be either stable or unstable. The latter type, known as radioisotopes, are radioactive.MASS:The amount of matter an object contains.MOLAR MASS:The mass, in grams, of1 mole of a given substance. The value in grams of molar mass is always equal to the value, in atomic mass units, of the average atomic mass of that substance. Thus carbon has a molar mass of 12.01 g, and anaverage atomic mass of 12.01 amu.MOLE:The SI fundamental unit for "amount of substance." A mole is, generally speaking, Avogadro's number of atoms or molecules; however, in the more precise SI definition, a mole is equal to the number of carbon atoms in 12.01 g of carbon.MOLECULE:A group of atoms, usually, but not always, representing more than one element, joined in a structure. Compounds are typically made of up molecules.PERIODIC TABLE OF ELEMENTS:A chart that shows the elements arranged in order of atomic number, along with chemical symbol and the average atomic mass (in atomic mass units) for that particular element.RADIOACTIVITY:A term describing a phenomenon whereby certain isotopes known as radioisotopes are subject to a form of decay brought about by the emission of high-energy particles. "Decay" does not mean that the isotope "rots"; rather, it decays to form another isotope, until eventually (though this may take a long time), it becomes stable. views updated Jun 11 2018The atomic mass of an atom is the mass of that atom compared to some standard, such as the mass of a particular type of carbon atom. The terms atomic mass and atomic weight are often used interchangeably, although, strictly speaking, they do not mean the same thing. Mass is a measure of the total amount of matter in an object. Weight is a measure of the heaviness of an object. In general, the term atomic mass is preferred over atomic weight.Scientists usually do not refer to the actual mass of an atom in units with which we are familiar (units such as grams and milligrams). The reason is that the numbers needed are so small. The mass of a single atom of oxygen-16, for example, is 2.657 × 10−23 grams, or 0.000 000 000 000 000 000 000 026 57 grams. Working with numbers of this magnitude would be very tedious.HistoryEarly chemists knew that atoms were very small but had no way of actually finding their mass. They realized, however, that it was possible to express the relative mass of any two atoms. The logic was as follows: suppose we know that one atom of hydrogen combines with one atom of oxygen in a chemical reaction. It is easy enough to find the actual masses of hydrogen and oxygen that combine in such a reaction. Research shows that 8 grams of oxygen combine with 1 gram of hydrogen. It follows, then, that each atom of oxygen has a mass eight times that of a hydrogen atom.This reasoning led to the first table of atomic masses, published by John Dalton (1766–1844) in 1808. Dalton chose hydrogen to be the standard for his table of atomic masses and gave the hydrogen atom a mass of 1. Of course, he could have chosen any other element and any other value for its atomic mass. But hydrogen was the lightest of the elements and 1 is the easiest number for making comparisons.One problem with which Dalton had to deal was that he had no way of knowing the ratio in which atoms combine with each other. Since there was no way to solve this problem during Dalton's time, he made the simplest possible assumption: that atoms combine with each other in one-to-one ratios (unless he had evidence for some other ratio).The table Dalton produced, then, was incorrect for two major reasons. First, he did not know the correct combining ratio of atoms in a chemical reaction. Second, the equipment used at the time to determine mass ratios was not very accurate. Still, his table was an important first step in determining atomic masses. Some of the values that he reported in that first table were: nitrogen: 4.2; carbon: 4.3; oxygen: 5.5; phosphorus: 7.2; and sulfur: 14.4.Words to KnowAtomic mass unit (amu): A unit used to express the mass of an atom equal to exactly one-twelfth the mass of a carbon-12 atom.Isotopes: Two or more forms of an element with the same atomic number (same number of protons in their nuclei), but different atomic masses (different numbers of neutrons in their nuclei).Mass: Measure of the total amount of matter in an object.Standard: A basis for comparison; with regard to atomic mass, the atom against which the mass of all other atoms is compared.Weight: The measure of the heaviness of an object.Within two decades, great progress had been made in resolving both of the problems that troubled Dalton in his first table of atomic masses. By 1828, Swedish chemist Jöns Jakob Berzelius (1779–1848) had published a list of atomic masses that was remarkably similar to values accepted today. Some of the values published by Berzelius (in comparison to today's values) are: nitrogen: 14.16 (14.01); carbon: 12.25 (12.01); oxygen: 16.00 (16.00); phosphorus: 31.38 (30.97); and sulfur: 32.19 (32.07).StandardsOne of the major changes in determining atomic masses has been the standard used for comparison. The choice of hydrogen made sense to Dalton, but it soon became clear that hydrogen was not the best element to use. After all, atomic masses are calculated by finding out the mass ratio of two elements when they combine with each other. And the one element that combines with more elements than any other is oxygen. So Berzelius and others trying to find the atomic mass of elements switched to oxygen as the standard for their atomic mass tables. Although they agreed on the element, they assigned it different values, ranging from 1 to 100. Before long, however, a value of 16.0000 for oxygen was chosen as the international standard.By the mid-twentieth century, another problem had become apparent. Scientists had found that the atoms of an element are not all identical with each other. Instead, various isotopes of an element differ slightly in their masses. If O = 16.0000 was the standard, scientists asked, did the 16.0000 stand for all isotopes of oxygen together, or only for one of them?In order to resolve this question, researchers agreed in 1961 to choose a new standard for atomic masses, the isotope of carbon known as carbon-12. Today, all tables of atomic masses are constructed on this basis, with the mass of any element, isotope, or subatomic particle being compared to the mass of one atom of carbon-12.Modern atomic mass tablesThe atomic mass of an element is seldom a whole number. The reason for this is that most elements consist of two or more isotopes, each of which has its own atomic mass. Copper, for example, has two naturally occurring isotopes: copper-63 and copper-65. These isotopes exist in different abundances. About 69.17 percent of copper is copper-63 and 30.83 percent is copper-65. The atomic mass of the element copper, then, is an average of these two isotopes that takes into account the relative abundance of each: 63.546.Students sometimes wonder what unit should be attached to the atomic mass of an element. For copper, should the atomic mass be represented as 63.546 g, 63.546 mg, or what? The answer is that atomic mass has no units at all. It is a relative number, showing how many times more massive the atoms of one element are compared to the atoms of the standard (carbon-12).Still, occasions arise when it would be useful to assign a unit to atomic masses. That procedure is acceptable provided that the same unit is always used for all atomic masses. Scient