Revisiting Molar Mass, Atomic Mass, and Mass Number: Organizing, Integrating, and Sequencing Fundamental Chemical Concepts

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The molar mass, atomic mass, and mass number of an element are valuable and fundamental pieces of information that chemists as well as students of chemistry use when solving problems. But for an introductory chemistry student, these terms can pose a formidable conceptual challenge if one goes beyond the simplistic definitions and tries to describe the differences and commonalities between the concepts. Sadly enough it is the articulation of these relationships that is missing in currently available chemistry textbooks for introductory students. Take for example some popular college textbooks *(1–5)* and one can see that while these three terms are defined early in the texts such as in chapters 1 through 3, the concepts are not integrated into any meaningful whole. Upon closer inspection, it is also apparent that the descriptions of these concepts often follow the same order: first mass number, followed by atomic mass, and then molar mass.

In this *Journal,* educators have attempted to clarify atomic mass and some of its related concepts as far back as the 1930s *(6).* Currently, educators use classroom activities and laboratory experiments *(7),* have re-examined definitions *(8),* and used analogies *(9)* to help teachers and their students understand ideas relating to mass. Others have investigated the historical roots of these concepts. One chemical educator, for instance, believes that the concept of equivalent weight is important in elucidating the arrangement of elements in the periodic table *(10).*

The Atomic Mass Conception: Articulating the Three Concepts

The advantage of articulating conceptual knowledge is to gain a deeper, more coherent understanding than one would if concepts were studied in isolation. Therefore, describing, integrating, and sequencing the above three concepts, into what will be referred to as the "atomic mass conception", could provide teachers as well as learners with a greater, more sophisticated understanding on which to draw upon when solving chemistry problems. While it is not the intention to test this hypothesis, it is this underlying assumption that provides the motivation for writing this article.

The greatest utility of the three concepts that comprise the atomic mass conception is to allow chemists to distinguish between the fundamental building blocks of matter, from one element to the next and from one atom to the next. Describing if one atom has a mass six times as much as another or counting the assorted number of protons and neutrons of a variety of atoms is in effect describing their differences, differences that can be used in theories to explain the chemical behavior of substances.

Molar Mass

The first of the three concepts to be discussed is molar mass. This concept represents a bridge between mass and amount; it provides chemists with a way to isolate by mass a desired number of atoms of an element or compound in order to perform quantitative experiments. Because mass and amount as expressed by grams and moles reflect quantities that are visible, the concept of molar mass can be characterized as relating to the macroscopic world. Molar mass is usually defined as the mass of 1 mole of a substance and its units are usually represented symbolically as kg/mol.

Some instructors and textbooks mention the term equivalent weight usually in a historical context *(10).* As the precursor to molar mass, equivalent weight is one of the early terms chemists gave to the study of the combining ratios of elements and is based on how much an element weighs relative to another element. Early methods of determining equivalent weight involved gravimetric analysis and quantitative studies of gases *(11).*

While inventors of equivalent weight did not use moles to express quantities, the notion of using a relative scale is carried over to other concepts associated to molar mass. For example, the terms: molecular weight, relative molecular mass, and relative molar mass all involve relative quantities. Relative molecular mass symbolized by M_r is defined by IUPAC as the "ratio of the mass of a molecule to the unified atomic mass unit" *(16),* while molecular weight in the *CRC Handbook of Chemistry and Physics* is described as "the ratio of the average mass per molecule or specified quantity of a substance to 1/12 of the mass of nuclide ^{12}C" *(17).* Because the unified atomic mass unit refers to the 1/12 the mass of carbon-12, both definitions have a common denominator and are therefore "relative" to a specific quantity.

In the context of IUPAC terminology, molecular weight, relative molecular mass, and relative molar mass are considered to be synonymous with each other, but, they are not synonymous with molar mass. This is evident in the units for these terms. For example, carbon's molecular weight, relative molecular mass, and relative molar mass are 12.01 and are dimensionless, while carbon's molar mass is 12.01 and has units of g/mol.

Formula weight is another term used by chemists and is conceptually similar to molecular weight, but with one key difference. Chemists usually use formula weight when referring to ionic instead of molecular compounds.

Atomic Mass

The utility of the concept of atomic mass is simply to numerically indicate the mass of an atom in its ground state

(12). An atom's mass is expressed in the non-SI unit of u, which refers to unified atomic mass unit (previously atomic mass unit). Since one unified atomic mass unit is defined as 1/12 mass of a carbon-12 atom, the atomic mass of ^{12}C is made equal to 12 u. The atomic mass of ^{12}C refers to the total mass of the protons, neutrons, and electrons that make up the atom. Depending on the level of precision, electrons are usually not counted since they weigh three orders of magnitude less than neutrons and protons (electrons are approximately 1/1800 the mass of a neutron or proton). One unified atomic mass unit is approximately equal to 1.66054×10^{-27} kg.

Closely associated to the concept of atomic mass is average atomic mass. The word "average" is used to relate the fact that the numerical value assigned to each element in the periodic table is a composite value, one that reflects the average and abundances of the atoms that compose a naturally occurring element. For example, natural carbon consists of 98.89% ^{12}C and 1.11% ^{13}C, and the value of carbon's average atomic mass listed in the periodic table is 12.01 u and represents an average of the all species of carbon atoms based on their masses as well as their amounts or abundances in nature *(13)*.

Because of the usefulness of the periodic table and the assumptions inherent in language, chemists often will use the term "atomic mass" when they are actually referring to the average atomic mass of an atom. But for teaching introductory chemistry students it is important to use words that describe one's intent. Therefore, it is a misconception to view 12.01 u as the mass of a single atom of carbon. The atomic mass of a single atom of carbon such as ^{12}C is 12 u not 12.01 u. The discrepancy in values between the average atomic mass and atomic mass is much greater when one examines other elements like bromine or silver.[1]

In the context of the above example involving the average atomic mass of carbon, the general name given to the different species of atoms in an element is called isotopes. The term "isotope" derives from two Greek words: *iso*, which means equal, and *topos,* which means place. Therefore all the different atomic variations of an element are found in the equal or same place on the periodic table. While the first isotope was discovered during investigations with radioactive substances in the early 20th century it has since been found that many elements have multiple isotopes. Multiple isotopes such as carbon have the same number of protons but a different numerical sum of protons and neutrons (mass number). The difference in mass number between the isotopes is only slight since the number of neutrons changes by only small quantities. It is another misconception to think that an atom is only an isotope if its number of protons is different from its number of neutrons. For example, ^{12}C has 6 neutrons and 6 protons, but, it is still an isotope of the element carbon.

As I previously stated, average atomic mass usually has units of u and is associated to isotopes. Other terms, such as relative atomic mass, relative nuclidic mass, and atomic weight have a different meaning from average atomic mass. For example, relative atomic mass symbolized by A_r refers to the "ratio of the average mass of the atom to the unified atomic mass unit" *(12)* and is dimensionless. This dimensionless relationship is similar to the one between molar mass and relative molecular mass elucidated earlier. IUPAC terminology considers relative atomic mass, relative nuclidic mass, and atomic weight to be synonymous with each other although they have different historical origins. For instance, atomic weight is the name Dalton used in the early 19th century to numerically describe the weight of atoms relative to each other. He drew upon known values of equivalent weight and his belief in atoms to create the first organized list of atoms by weight.

One of the most apparent differences between the first concept, molar mass, and the second concept, atomic mass, is the difference in scale. One moves from the macroscopic level of the gram and mole to subatomic particles of protons and neutrons expressed in u. Although this difference in scale exists, there is a mathematical relationship that ties these concepts together through the conversion of their units from grams per mole to u.

To understand this relationship let's begin by splitting apart the conversion. The numerator of the units, grams, is converted to u by utilizing the equality, 1 g = 6.02×10^{23} u. The generation of this equality is evident from eqs 1–3 where ^{12}C is used as an example. These equations rest on two definitions of a mole, one that relates to Avogadro's number and the second that relates to 12 g of carbon. For example, one textbook author describes the latter relation when he defines the mole as "the number equal to the number of carbon atoms in exactly 12 grams of pure ^{12}C" *(1).*

$$12\text{ g} = \left(6.02214 \times 10^{23}\text{atom}\right)\left(12\frac{\text{u}}{\text{atom}}\right) \tag{1}$$

$$12\text{ g} = \left(6.02214 \times 10^{23}\right)\left(12\text{ u}\right) \tag{2}$$

$$1\text{ g} = 6.02214 \times 10^{23}\text{ u} \tag{3}$$

Although it is not commonly written in texts, the units of u can be written as u per atom, not just solely as u. This is important because the denominator "mol" of molar mass must be converted to atoms. This conversion can be accomplished by using the definition of mole.

It is evident that two relationships exist between molar mass and average atomic mass. First, there is a mathematical relationship between two different sets of units and secondly, there is a numerical relationship that remains constant before and after the conversion. For example, in the latter relationship involving carbon the value of 12.01 does not change during the conversion of the units. This is because both molar mass and average atomic mass involve the mass of an atom or group of atoms (i.e., mole) relative to a number of atoms or group of atoms. An important difference is that average atomic mass involves isotopes, while molar mass does not; in other words molar mass lacks a conceptual counterpart to isotopes *(14)*.

The concepts related to atomic mass have been discussed in terms of non-nuclear chemical reactions. The introduction of nuclear chemistry, which usually occurs in later semesters of chemistry or in physics courses, takes into account the changing mass of the nucleus owing to energy changes, what is usually referred to as mass defect. This phenomenon is usually not detectable with normal chemical reactions and therefore can be assumed to be negligible *(1)*.

Mass Number

Mass number is synonymous with the term nucleon number and refers to the total number of protons and neutrons in the nucleus of an atom. The symbol for the quantity is *A* and is dimensionless. The utility of this concept lies with providing scientists a way to distinguish between isotopes.

The concept of mass number is tied to the concept of atomic number. Both concepts were introduced during the early parts of the 20th century when radioactive substances were being studied and the parts of the atom were being described. Concepts of mass number, along with atomic number, were used by scientists to keep track of atomic particles that change during radioactive decay. In the following equation, ${}^{14}_{6}C \rightarrow {}^{14}_{7}N + {}^{0}_{-1}e$, it is evident that ^{14}C disintegrates into elemental nitrogen by losing an electron beta particle. The nucleus of carbon decays from six protons and eight neutrons to a nucleus representing nitrogen having seven protons and seven neutrons. The concepts of mass and atomic numbers allow a scientist to not only represent changes in the nucleus, but how many of the *total* neutrons and protons of an atom change.

The atomic number rather than the mass number became a basis for ordering elements in the periodic table. This is because the number of protons did not change from one isotope to another, instead, it was found that the number of protons changed by increments of one from element to element. Once the concepts of atomic and mass numbers were developed chemists no longer preferred the organization of the elements by mass. Nevertheless, mass was not totally disregarded as evident by the discussion of the two earlier concepts, molar mass and the average atomic mass, and their numerical appearance in the periodic table.

The link between mass and the number of nuclear particles was experimentally deduced by Rutherford. Determining the charge on an atom's nucleus led to the calculation of the number of electrons; this value was found to be about half the atomic weight of the element. This meant that for an electrically neutral atom, the number of electrons and protons were about equal to an element's atomic weight or in present terms, average atomic mass *(11).*

Because of the almost even relationships between the number of protons, neutrons and electrons of an atom, the number of particles that is represented by the mass number is close to the average atomic mass of an atom. This is possible because of the use of a relative scale to calculate an atom's mass. It is reasonable to believe that if a carbon atom is about twelve times heavier than a hydrogen atom, then carbon would have about twelve times the number of particles as hydrogen. While this is true, this regularity also can be misleading if larger atoms are considered. Therefore, simply put, there is a connection between mass number and average atomic mass but it is not an exact one.

The Creation of New Problems

Problems that introductory chemistry students encounter involving molar mass, atomic mass, and mass number are similar from textbook to textbook. These questions are often mathematically simplistic and rely mostly on operational definitions. In contrast to this, the subject of this article has provided a basis for asking students a variety of questions that involve a much more sophisticated understanding of these concepts. The following questions listed below are just a few examples that would be appropriate on the introductory level.

1. As the name implies, the average atomic mass of an atom is based on an average of isotopes. Is the molar mass of an element also based on an average of elements? Explain.

   Answer: No. Multiple carbon elements do not exist; there are only multiple carbon atoms called isotopes.

2. What numerical value relates molar mass and average atomic mass?

   Answer: Avogadro's Number.

3. What is a key difference between molar mass and atomic mass in terms of scale?

   Answer: Molar mass relates to the macroscopic level, while atomic mass relates to the atomic level.

4. Is the mass number of carbon 12.01? Explain.

   Answer: No. The mass number refers to a specific isotope of carbon. The value of 12.01 refers to an average based on abundances of naturally occurring carbon isotopes.

5. The mass number of ^{12}C is 12, what is the average atomic mass of carbon naturally found on earth?

   Answer: 12.01 u. Notice the absence of a superscript and the addition of units when moving between the concepts of mass number and average atomic mass.

6. Why is the unit u used instead of kilograms when discussing the average atomic mass of an atom?

   Answer: Kilograms could be used, but it is a matter of convention. Also if mass were used the obvious connection to the number of protons and neutrons would be lost. The invention of the unified atomic mass unit underscores the proton and neutron relationship.

7. Name two key ideas involved when converting from g/mol to u.

   Answer: The two ideas can be expressed mathematically: (i) 1 mol = 6.02×10^{23} atoms and (ii) 1 g = 6.02×10^{23} u.

8. Is it true that 1 g is equal to 1 u? Explain.

   Answer: It is a misconception to think that 1 g is equal to 1 u even though 12 g/mol can be converted to 12 u/atom. Rather, 1 g is equal to 6.02×10^{23} u.

9. What is the difference between molar mass and relative molecular mass?

   Molar mass is the mass of 1 mol of a substance usually expressed in g/mol. Relative molecular mass has the same value of molar mass but has a unit of 1. This is because relative molecular mass by definition is a ratio of the mass of a molecule to the unified atomic mass unit.

Sequencing Issues

An important issue for educators to consider is the presentation of these concepts in textbooks as well as in the classroom. Should molar mass precede the introduction of atomic mass, or should it follow from its discussion? Does a specific sequence make more sense to teachers as well as students? In introductory chemistry textbooks mass number usually precedes atomic mass, which in turn precedes the concept of molar mass *(1–5).* The textbook order of these concepts moves in a discussion dominated by an atomic level description of matter to molar quantities on the macroscopic level. One might hypothesize that the reason textbook writers order these concepts in the way they do involves mathematics. The math associated to the three concepts increases in sophistication of problem solving: the concept of mass number involves simple arithmetic while the concept of molar mass is used in stoichiometry problems when mass and moles are converted.

The sequence presented in this article—molar mass followed by atomic mass and then by mass number—is the reverse of the usual textbook order of concepts. The author's sequence reflects historical developments of these ideas. For instance, molar mass is associated to the early macroscopic investigations of equivalent weight, which precedes in time the concept of atomic mass and the development of the atom, which then precedes mass number and the description of the nucleus. Introducing molar mass as the first of the three concepts is to underscore the importance chemists placed on differentiating and characterizing elements. Starting with mass number, as textbooks often do, overlooks the importance of empirical investigations and diverts attention to matters of isotopes. An introduction of the concept of isotopes only really makes sense if atomic mass is first understood.

Additional support for presenting molar mass before mass number is found when the role of theory and explanation are considered. If one believes that the *utility of the concept of atoms is to explain observable matter rather than matter used to explain atomic behavior*, then, one can create an argument that places the observation of matter before its explanation. Beginning a discussion with molar mass would place empirical observations and measurements before atomic explanations.

While a conceptual order informed by history can bring a greater sense of authenticity to the classroom, a larger more problematic question arises. Do young adult learners of chemistry learn better when instruction centering on a macroscopic discussion of matter precedes discussions of the atomic level? This problem takes on increasing complexity when one considers the types of learners involved, the instructional methods used, as well as contextual factors. Be this as it may, educational theory has much to say on the importance of a learner's sensory experiences in the classroom. According to Saunders, direct sensory experiences with macroscopic objects such as investigating chemical reactions in the laboratory provide two important opportunities for learners *(15).* The first is sensory experiences can involve actual phenomena or materials that scientists have or are currently using or studying, thereby allowing students to participate in a more realistic experience of science. Secondly, sensory experiences can provide learners with an opportunity to modify their understanding in lieu of their observations. This latter point is a critical aspect of learning considering the large number of documented misconceptions in chemistry that introductory students possess *(16).*

The importance of what learners know and experience, their prior knowledge, has persuaded Fensham to recommend that teachers should begin the teaching of chemistry in terms of macroscopic chemical reactions rather than the structure of atoms *(17).* Fensham bases his position on research studies that have repeatedly shown that students, even those with extensive study of the kinetic molecular theory and of atomic and molecular structure, give macroscopic rather than atomic explanations for chemical phenomena. Fensham believes that since students have this strongly held perspective it would be appropriate for them to begin chemistry with discussions related to the macroscopic level.

As the question has turned from why molar mass should precede mass number to why discussions of the macroscopic level should precede the atomic level, it is apparent that sequencing of concepts must not just be based on a teacher's logical understanding of a subject. Rather one should consider what the learner knows, educational research, the contextual and historical factors involved in the invention of the concept as well as the type of relationships that are to be discussed between concepts. The relationships between molar mass, atomic mass, and mass number have been articulated to provide chemical educators with a nuanced description that goes beyond an introductory textbook treatment of these fundamental concepts.

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Note

1. Likewise, when mass spectroscopy is used to analyze high molecular compounds such as bovine insulin having the formula $C_{254}H_{377}N_{65}O_{75}S_6$, there is a significant difference between the lowest calculated or nominal mass, the monoisotopic mass, and average mass of the compound *(18).* IUPAC defines a monoisotopic mass spectrum as a spectrum "…containing only ions made up of the principal isotopes of atoms making up the original molecule." Taken from the IUPAC Web page; see ref *12* for exact source.

Literature Cited

1. Zumdahl, S. S.; Zumdahl, S. *Chemistry,* 5th ed.; D. C. Heath: Lexington, MA, 2000.
2. Ebbing D. D.; Gammon, S. D. *General Chemistry,* 6th ed.; Houghton Mifflin: New York, 2000.
3. Brown, T. L.; LeMay, H. E.; Bursten, B.; Burdge, J.; Brown, T. E. *Chemistry: The Central Science,* 9th ed.; Prentice Hall: Upper Saddle River, NJ, 2002.
4. Chang, R. *Chemistry;* McGraw Hill: New York, 2002.
5. Silberberg, M. *Chemistry: The Molecular Nature of Matter and Change,* 3rd ed.; McGraw Hill: NY, 2002.
6. Wildman, E. A. *J. Chem. Educ.* **1933,** *10,* 238.
7. Horsch, E.; Burnett, D. *J. Chem. Educ.* **1995,** *72,* 435. Huff, R. B.; Evans, D. W. *J. Chem. Educ.* **1991,** *68,* 675. Bodner, G. M.; Magginnis, L. J. *J. Chem. Educ.* **1985,** *62,* 434.
8. Richardson, D. E. *J. Chem. Educ.* **1992,** *69,* 736. Baumgartner, E.; Benitez, C.; Cirelli, A. F.; Flores, L. L. *J. Chem. Educ.* **1988,** *65,* 16.
9. deSanabia, J. A. *J. Chem. Educ.* **1993,** *70,* 233. Last, A. M.; Webb, M. J. *J. Chem. Educ.* **1993,** *70,* 234. Fortman, J. J. *J. Chem. Educ.* **1993,** *70,* 235.

10. Laing, M. *J. Chem. Educ*. **1996**, *73,* 1007.
11. Leicester, H. M. *The Historical Background of Chemistry;* Dover Publications: New York, 1971.
12. International Union of Pure and Applied Chemistry. *IUPAC Compendium of Chemical Terminology*: *http://www.iupac.org/publications/compendium/index.html* (accessed Jan 2005).
13. *CRC Handbook of Chemistry and Physics*, 3rd electronic ed.; Lide, D., Ed.; CRC Press: Boca Raton, FL, 2004.
14. Herron, J. D. *The Chemistry Classroom: Formulas for Successful Teaching;* American Chemical Society: Washington DC, 1996.
15. Saunders, W. L. *School Science and Mathematics* **1992,** *92,* 136–141.
16. Birk, J. P.; Kurtz, Martha J. *J. Chem. Educ.* **1999**, *76,* 124. Sanger, M. J.; Greenbowe, T. J. *J. Chem. Educ.* **1997**, *74,* 819. Beall, H. *J. Chem. Educ*. **1994,** *71,* 1056. Nakhleh, M. B. *J. Chem. Educ*. **1992,** *69,* 191.
17. Fensham, P. Beginning To Teach Chemistry. In *The Content of Science: A Constructivist Approach to its Learning;* Fensham, P., Gunstone, R., White, R., Eds.; Falmer Press: London: 1994.
18. Yergey, J. A. *International Journal of Mass Spectrometry and Ion Physics* **1983,** *52,* 348.