Why Do Chemists Weigh Things in Grams Instead of Gamus?

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Abstract: This paper points out that chemists still call the determination of mass "weighing," when doing so is incorrect. Legitimizing mass as a verb is advocated. It then suggests a new unit of mass for chemists, the gamu, or grand amu, that will change Avogadro's number to an even ten to the 24th power. This will improve student understanding of macroscopic to microscopic conversions, while at the same time making these conversions computationally easier. Examples are given to compare calculations using both gamus and grams.

First, they don't weigh things in grams, and second they don't use gamus (pronounced gam'-oos) because they don't know what a logical and handy unit it is. Allow me to explain.

Physicists quit weighing things in grams or kilograms years ago. They mass things. Kilograms and grams are units of mass, not units of weight. Weight is a force, and forces are measured in dynes (old physicists) or newtons (young physicists). The balances scientists use do not weigh things. They balance the mass of the object in question with some known masses and infer the mass of the sample from that experiment. Weight is involved in this process, but cancels out; therefore, the balance instruments are, to a degree, independent of altitude and give the same result in Bogalusa, Louisiana and in Ouray, Colorado. Because the whole process of using a balance involves balancing two weights, chemists in orbit who want to mass things traditionally radio in and say "Houston. We have a problem."

For some reason difficult to understand, chemists still say, "weigh out a sample" rather than "mass out a sample." Chemists are using the balance principle that cancels the weights, so they are massing their samples. When they are through with the operation, they do not know the weight of the object in question. Yet, they still say, "weigh," as if they were using spring scales (which do measure the force). Don't try to get chemists to change; they probably won't [1–3]. Chemists may forever say, "weigh out a sample," though they know full well that they do not want the student to come back and say, "This sample weighs 12.3 millinewtons in this geographic location."

Students catch on quickly. They say, "This sample weighs 1.256 grams." They don't say "This sample masses 1.256 grams" because their teachers don't. The use of weigh (determine the force) instead of determine the mass is confusing only to the students who just had beginning physics. Some day chemists will have to face the fact that, on balance, this is not the weigh to go [4].

Now for the second question. Why do chemists not mass things in gamus? I really don't know a good answer to this one. A very likely answer is that I haven't told enough of them about gamus yet. Otherwise, why would chemists **not** use such a logical and simple unit of mass, which totally un-confuses the students about the distinctions between the macroscopic and the microscopic worlds of chemistry? The answer requires that the reader be familiar with a small mass unit called the amu (atomic mass unit), the gram (an arbitrary mass unit), and the gamu (grand atomic mass unit). Also the reader needs to appreciate that Avogadro's number is simply the conversion factor between amus and grams. There are 6.02×10^{23} amus per gram.

I digress here briefly to tell you a story. This happened to me when I was a chemistry graduate student at a large university. I was sharing an office with one of the department's famous inorganic chemists (who will remain anonymous) when a student of Dr. X's came in to ask a very good question. "Dr. X. You told us that recently they changed the standard of atomic masses from O (natural isotopic mixture) equals exactly 16 amu to carbon 12 equals exactly 12 amu. What I want to know is, did that change Avogadro's number?" After brief thought Dr. X said, "No, it did not."

Tactlessly, I (a mere graduate student) butted in and said "Excuse me Dr. X, but since Avogadro's number is the number of amus per gram, and they just changed the size of the amu, but not the size of the gram, then this redefinition of the amu **did** change Avogadro's number, though very slightly." It says a lot for Dr. X that he was not angry at this correction, but thought about it calmly and finally agreed that I was right and he was wrong.

So, back to the concept of there is Avogadro's number of amus per gram. The gamu, short for the **grand amu**, I hereby define as a trillion trillion amus. 1×10^{24} amus equal 1 gamu; therefore, if in your lab, you mass out 24.31 gamus of magnesium you obviously have massed out a trillion trillion atoms of magnesium. If you then oxidize this magnesium (close your eyes during this process) and find that the resulting compound masses out at 16 gamus bigger than the starting material, it isn't such a big mental hurdle to see that if 24.31 gamus of Mg combined with 16 gamus of O, then a trillion trillion Mgs combined with a trillion trillion Os to give a trillion trillion MgOs, or that there is one O for each Mg.

The same experiment done in grams would, of course, give the same ratio of masses, but the number of atoms that reacted and the number of MgO pairs formed would not be the large **but comprehensible** number trillion trillion, but the **less comprehensible** number 6.02×10^{23} . In other words, by defining a new mass unit (the gamu) we change the equivalent of Avogadro's number from 6.02×10^{23} to 10^{24}, a trillion trillion.

And so, just as Avogadro's number is 6.02×10^{23} things, our new number of things is 10^{24}. Because a mole is Avogadro's number of things, then 10^{24} things cannot also be called a mole. We shall call it a mule. If chemists mass out their reagent in mules, it is much simpler to see that this is a macroscopic mass exactly 10^{24} times larger than what goes on at the atomic and molecular level.

If you are confused as to how this makes things simpler for the students, do this: show your students a beaker containing 180 grams of water and ask them to tell you how many atoms are in this sample. If you allow calculators, some of the students will arrive at the correct solution. Now, hold up a beaker containing 180 gamus of water and ask how many atoms are there. No calculators are necessary. Obviously, this is ten mules, or 10^{25} molecules times 3 atoms per molecule equals 3×10^{25} atoms.

Now ask the class to calculate the mass of a single carbon dioxide molecule in grams and in gamus. Again, they will need calculators for the first job. The answer to the second question is 44×10^{-24} gamus, of course.

So, how much larger is a gamu than a gram? It is 1×10^{24} divided by 6.02×10^{23}, which is approximately 1.66 times larger. Can we persuade balance manufacturers to provide a gamu scale on the balances? With modern electronic balances this would be no problem. Of course both gram and gamu readout would be available. If a sample of chromium masses at 52 milligamus we know immediately this is 1×10^{21} atoms of chromium. How convenient. Is this a millimole of chromium? No. It is too large to be a millimole. It is a millimule of chromium. Mules are larger than moles, as everyone knows.

I know what you are thinking. If we change the standard from grams to gamus will it not change important constants like *R* and k? Well let us see. *R*, the gas law constant, has units of joules per (mole Kelvin). If we wish to change it to joules per (mule kelvin) we need only multiply by 1.66 which gives a gas law constant of 13.8. Now the gas law constant per molecule (Boltzman's constant) will be 13.8×10^{-24}, because 10^{24} is the number of molecules in a mule. No need to remember two numbers any more. The values of *R* and *k* differ only in their power of ten. To take an example of using the new *R* value:

What is the pressure exerted by the gas in a container of volume one cubic meter if the gas is 32 gamus of oxygen at 25 degrees C?

ANSWER: $P = nRT/V = \{1 \text{ mule} \times 13.8 \text{ joules/mule/Kelvin} \times (25 + 273) \text{ Kelvin}\}/1 \text{ m}^3 = 4110 \text{ N/m}^2$.

OLD ANSWER: $P = nRT/V = \{[32 \times 1.66] \text{ g}\}/32\text{g mol}^{-1}\} \times 8.314 \text{ J mol}^{-1} \text{ K}^{-1} \times (25 + 273) \text{ Kelvin}\}/1\text{m}^3 = 4110 \text{ N/m}^2$.

In a similar fashion $R = 0.082$ L atm mol^{-1} K^{1}) becomes $0.082 \times 1.66 = 0.136$ L atm mule^{-1} K^{-1}). The famous 22.4 L mol^{-1} at STP becomes 37.2 L mule^{-1} at the same conditions, because this volume contains 10^{24} molecules, not 6.02×10^{23} molecules.

This probably seems confusing. Is this really going to help anybody understand chemistry? Yes, it will help students. It will not help teachers who grew up on moles. If a student learns mules instead of moles they need not learn Avogadro's big magical number. A mule of things will become as simple a concept as 100 things, except the number is larger. What really is happening here is that, becuase I cannot change the mass of atoms, and because the number relating these small masses to visible large masses depends upon the arbitrary unit of mass that was adopted before Avogadro (the gram), then I propose changing the arbitrary unit of mass. Grams are not sacred, though admittedly quite ingrained in the culture. Gamus, symbolized gm, could catch on, and chemistry students the world over would be grateful. When that happens, the world we see can be said to be a trillion trillion times as large as the world we symbolize on our blackboards.

Which reminds me: If blackboards can disappear, replaced by something more convenient, then so can grams. Goodbye grams. Hello gamus.

References and Notes

1. Plumsky, R. *J. Chem. Educ.* **1996,** *73,* 451–454.
2. Smith, M. W.; Russell, C. E. *J. Chem. Educ.* **1997,** *74,* 900.
3. What do the Books Say? Actually, authors of general chemistry texts are trying to change. They now use the term molecular mass instead of molecular weight. They say, "The mass of a sample of…is…." They almost always have a brief paragraph stating that weight and mass is not the same thing. Most chemists, however, still tell students to "Weigh the beaker, then the beaker with sample, and subtract the weights." There are even cases of "Determine the mass of the sample and container and then subtract the weight of the container." This definitely cannot be done.

   Quotations From Some Texts.

   Caution: The SI units for mass and weight are often misused in everyday life. Incorrect expressions such as "This box weighs 6 kg" are nearly universal. What is meant is that the mass of the box, probably determined indirectly by weighing, is 6 kg. This usage is so common that there is probably no hope of straightening things out, but be sure you recognize that the term weight is often used when mass is meant. Be careful to avoid this kind of mistake in your own work! In SI units weight (a force) is measured in newtons, while mass is measured in kilograms [Young, H. D.; Freedman, R. A. *Sears and Zemansky's University Physics,* 10 ed.; Addison-Wesley: Reading, MA, 2000; p106].

   Chemists are interested primarily in mass, which can be determined readily with a balance; the process of measuring mass, oddly, is called *weighing*. [Chang, R. *Chemistry,* 7 ed.; McGraw-Hill: New York, 2002, p. 15].

   Strictly speaking, the pound is a unit of weight rather than mass. The weight of an object depends on the local force of gravity. For measurements made at the earth's surface the distinction between mass and weight is not generally useful. [Moore, J. W.; Stanitski, C. L.; Jurs, P. C. Chemistry: The Molecular Science; Harcourt College Publishers: Orlando, FL, 2002, p 46].

   Because weighing something on a chemical balance (see Fig. 1.5) involves comparing the mass of that object to a standard mass, the terms weight and mass are sometimes used interchangeably, although this is incorrect. [Zumdahl, S, S.; Zumdahl, S. A. Chemistry, 5th ed., Houghton-Mifflin: Boston, MA, 2000, p 9].

   The object on a balance pushes the pan down with a force equal to m × g, where m is the mass of the object and g is the acceleration of gravity. The electronic balance uses an electromagnetic restoring force to return the pan to its original position. The electric current required to generate the force is proportional to the mass, which is displayed on a digital readout. [Harris, D. C. Quantitative Chemical Analysis, 5th ed.; W. H. Freeman: New York. 1999, p. 32.

4. Let's Make Mass a Verb. One could say, "why should we care whether the process is called weighing or massing?" There are several reasons:

   1. The student is taught something that is demonstrably incorrect. Namely, that one can think of weight and mass as the same thing. One surely cannot, for they don't even have the same units. All the examples about "mass is the same on Mars as on Earth, but the

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weight is different" seem to suggest to the student that on Earth the mass and the weight are well behaved, that is, equal to each other.

2. The statement that the observed weight = $g \times m$ is not really true. A more correct statement is observed weight ± buoyancy correction = (g ± location correction) × m. The location correction is taken care of by calibration with standard masses, commonly called standard weights. The buoyancy correction is technically required in all cases where the density of the calibration masses is not the same as the density of the sample; however, this buoyancy correction is often ignored except when massing gases. Buoyancy corrections are not easy to explain to a student who has adopted the idea that weight and mass are interchangeable.

3. How can one be said to **weigh** a sample when, after the operation is concluded, one does not know the weight?

4. I cannot think of any other measurement we commonly make that we do not say, "measure the (current, volume, pH, conductivity, etc.)," rather than use one operation name like "weigh." If we refuse to say measure the mass, we should make mass a verb, and say mass your sample.