Commentary

A Plea for the Abandonment of the Atmosphere As a Unit in Gas Law Instruction

by George C. McBane

Few quantities cause so much trouble with units as does pressure. In a typical general chemistry course, the student encounters pressure measured in atmospheres, pascals, Torr, and pounds per square inch. The conversion factors among all these units are apparently arbitrary and follow no sensible pattern. In addition the gas constant, R, appears with quite different numerical values when it is expressed using each of those pressure units. (Thankfully, the version of R using pounds per square inch is rarely used.) R also occurs in general chemistry in energy units, as 8.314 J/(mol K). In my experience, students find the various numerical values of R mysterious, can see no connection between the energy and pressure–volume versions, and usually must resort to simple unit analysis to decide which particular value of R is needed in a given problem.

In most parts of general chemistry, textbooks and instructors are fairly good about using consistent metric units and adopting appropriate units from the SI for use by students. For example, we have now made the switch from the calorie to the joule as the principal energy unit for thermodynamics. The atmosphere, however, has somehow clung to life as a pressure unit despite its disconnection with SI. In the United States, most general chemistry texts continue to describe elementary gas law calculations using the atmosphere as a pressure unit and using $R = 0.08206$ L atm/(mol K). I checked nine recent textbooks (1–9), which surely constitute a representative sample of current U.S. college texts. All nine use the atmosphere as the primary unit in gas law descriptions and sample calculations.

The base SI unit of pressure, the pascal (1 Pa = 1 N/m^2 = 1 kg m^{-1}s^{-2}), is probably not popular for general chemistry gas problems because it is too small a unit: pressures near ordinary atmospheric pressure have values around 10^5 Pa. However, two units, the kilopascal and the bar, have convenient sizes and mesh well with other SI units. The bar (1 bar = 10^5 Pa) is not strictly part of SI but is officially tolerated: it is described in the SI documentation (10) as "used in particular circumstances" and "likely to continue to be used for many years". Ordinary atmospheric pressure is near 100 kPa or 1 bar.

We chemistry educators should abandon the atmosphere and replace it with the bar or the kilopascal or both. One reason to do so is that in general it is better to teach science with a consistent unit system. But there is a stronger, pedagogical reason: when good pressure units are used in gas law problems, the physical equivalence between the pressure–volume and energy versions of the gas constant R becomes much more apparent. Simple unit analysis indicates that 1 Pa·1 m^3 = 1 J, so that 1 kPa·1 L = 1 J and 1 bar·1 L = 100 J. The gas constant is therefore $R = 8.314$ J/(mol K) = 8.314 kPa L/(mol K) = 0.08314 bar L/(mol K). Given that set of values, students have no trouble recognizing that there is simply a set of unit conversions going on, rather than different "versions" of the gas constant appropriate for different contexts.

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More fundamentally, the SI pressure units clarify the physical relation between the pressure–volume unit and energy unit versions of R (11). A pressure is a force divided by an area, or force divided by a distance squared, as in N/m^2; a volume is a distance cubed, m^3. So it makes sense that a pressure times a volume is a force times a distance, N·m, which students recognize as representative of work, with units of energy.

Both the bar and the kilopascal are convenient units for constructing simple, realistic problems for application of gas laws. Bourdon gauges of the sort used on gas tank regulators often have scales marked both in pounds per square inch and in kilopascals. 100 kPa or 1 bar is just as good a typical atmospheric pressure for sample problems as 1 atm, and for most places above sea level it's a better one. So there's no "convenience cost" for making the switch, as there was in the change from calories to joules.

If we abandon the atmosphere, do we lose the beloved "one mole of ideal gas occupies 22.4 L at STP"? 1 bar has been the recommended thermodynamic standard state pressure since 1982 (12), and the IUPAC definition of STP is 0 °C and 100 kPa (13). The nine textbooks cited above all give the pre-1982, obsolete definition of STP. At 1 bar and 0 °C, one mole of ideal gas occupies 22.7 L. That's no worse than 22.4, and it comes with a big pedagogical gain.

Several different versions of "standard temperature and pressure", STP, are used as reference conditions for computing quantities of gases bought or sold commercially. But commercial operations have never been particularly good guides for scientific teaching or practice. I see no pedagogical value in retaining the scientifically obsolete 1-atm STP.

So please, textbook authors and chemistry teachers, at both high school and college levels: let's abandon the atmosphere as a pressure unit. The switch to better units will be much less difficult than the switch from calories to joules was. The obsolete atmosphere is making our jobs, and the jobs of the students, more difficult. It has no advantages other than tradition, and we are paying a high price in clarity for maintaining the tradition.

Acknowledgment

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