

1.6: The variables of state, pressure units, and the ideal gas law

Temperature is only one of the key ways that we can describe the properties of a gas. Counting up the number of molecules - or, more reasonably, the number of *moles* - is another.

These are two of the four conventional *variables of state* (or *state variable*) for a gas. The variables of state are considered the ways of describing the *entire* condition of an ideal gas.

State variable	Symbol	SI Unit (mks)	Abbreviation	Base SI units
Temperature	T	Kelvin	K	K
Count of substance	n	moles	mol	mol
Volume	V	cubic meters	m^3	m^3
Pressure	P	Pascals	Pa	$\text{kg m}^{-1} \text{s}^{-2}$ (or N m^{-2})

Volume units are very straightforward; they're length units in three dimensions, in keeping with a volume being proportional to any length dimension of an object cubed. Of course, the base SI units for volume, cubic *meters*, are *very* large indeed; nobody would actually want to lift a cube of anything that was a meter of length on each side. This is why most small-scale exercises will set units of volume to cubic *centimeters*, or the size of a cube 1 centimeter on each side; this is much more manageable.

Many modern physical chemistry textbooks will make mention of a cubic *decimeter*, the size of a cube 1 decimeter (or 10 centimeters) on each side. 1 dm^3 is just another way to describe one *liter*, which is the common metric unit of volume. Because of the cubing of 10, there are 1000 liters in a cubic meter; there are 1000 cubic centimeters in a liter, which is why the cubic centimeter and the milliliter are frequently equated with one another.

$$1 \text{ L} = 1 \text{ dm}^3 = 1000 \text{ cm}^3 = 1000 \text{ mL} \quad 1 \text{ m}^3 = 1000 \text{ dm}^3 = 1000 \text{ L}$$

Pressure units are much more plentiful, and much more convoluted. *Pressure* is defined simply as a force per unit area. The base SI unit of pressure is the *Pascal*, which is equivalent to SI measure of force (the Newton) divided by the SI measure of surface area (the square meter). This is a *very* small unit of pressure, approximately one-one hundred thousandth normal atmospheric pressure.

Historically standard atmospheric pressure has been defined by the height of a column of mercury that is supported by that pressure, in the nature of an experiment designed by one of Galileo Galilei's apprentices, Evangelista Torricelli. In his honor, a height of 1 millimeter of mercury in a barometric column is called a *torr*; standard atmospheric pressure, also known as *one atmosphere*, is equivalent to a height of 760 mm of mercury in this kind of column; this turns out to be about 101,300 Pascals (or 101.3 kiloPascals).

Most chemists use the atmosphere as their reference atmospheric pressure, but some prefer a truly metric unit, the *bar*, which is set to exactly 100,000 Pascals (exactly 100 kPa). Meteorological pressure is frequently communicated in *millibars*.

$$1 \text{ atm} = 760 \text{ torr} = 1.013 \text{ bar} = 101\,300 \text{ Pa} \quad 1 \text{ bar} = 750.2 \text{ torr} = 0.9872 \text{ atm} = 100\,000 \text{ Pa}$$

For reasons that are purely historic, we will prefer the atmosphere as our units of pressure - but there will be times when any of the four will be preferable. Vapor pressures are far more frequently communicated in torr (which, again, is the same thing as mm Hg). When we do computations in base-SI units, we will need Pascals; and again, bars are metric derivations from the Pascal. Atmospheres are simply most frequently used and most easily explained - a multiple of normal atmospheric pressure.

If we take the kinetic theory's two conditions of *ideality* of gases - the essentially zero volume of the gas molecules, and the lack of attraction or repulsion between gas molecules - as valid, we find that the ratio of pressure times volume of an ideal gas to number of moles times temperature is a *constant*:

$$R = \frac{PV}{nT} = 0.082058 \text{ L atm mol}^{-1} \text{K}^{-1} = 8.31447 \text{ J mol}^{-1} \text{K}^{-1}$$

This is the equation more commonly written as $PV = nRT$, the ideal gas law. R , like the value of k_B , is a constant that will turn up every now and again in our computations. Like the definition of temperature, we *can* derive this law from first physical principles; however, this is a familiar enough statement to us that we will simply take it as something we know, as a straightforward starting point.

There's one additional thing to note about these two constants we've worked with, the value of k_B and the value of R . You see a hint of that thing in the exponent on the value of k_B , the exponent of 10^{-23} . Divide the value of R by the value of k_B and make a note of what comes out:

$$\frac{R}{k_B} = \frac{8.31447 \text{ J mol}^{-1} \text{ K}^{-1}}{1.38065 \times 10^{-23} \text{ J K}^{-1}} = 6.0221 \times 10^{23} \text{ mol}^{-1} \rightarrow \frac{R}{k_B} = N_A$$

R and k_B are, in many ways, the exact same constant. R , on the one hand, deals with an energy per *mole* per Kelvin; k_B deals with an energy *per molecule* per Kelvin. The ratio of R to k_B is the familiar value of Avogadro's number; this connection can prove useful for us as we move forward.

This page titled [1.6: The variables of state, pressure units, and the ideal gas law](#) is shared under a [not declared](#) license and was authored, remixed, and/or curated by [Chuck Pearson](#).

- [1.3: The variables of state, pressure units, and the ideal gas law](#) by [Chuck Pearson](#) has no license indicated.