IDEAL GAS LAW

The ideal gas law is

\[ p V = n R_u T = \frac{m}{M} R_u T = m R T = N k T \]  \hspace{3cm} (1)

where

- \( p \) = the absolute pressure of the gas,
- \( V \) = the volume occupied by the gas,
- \( n \) = the number of moles of the gas,
- \( R_u \) = the universal gas constant (8.314 J/°K·g-mole = 1545 ft·lb/°R·lb-mole),
- \( T \) = the absolute temperature of the gas (e.g., °K),
- \( m \) = the mass of the gas,
- \( M \) = the molecular weight of the gas,
- \( R \) = the gas constant for a particular gas (=\( R_u/M \)),
- \( N \) = the number of gas molecules, and
- \( k \) = Boltzmann’s constant (1.381×10**-23 J/°K).

**Example 1**

Determine the volume occupied by 1 g-mole of an ideal gas at standard temperature and pressure (STP). Scientific STP is defined as 0°C and 1 atm.† Using the ideal gas law and the fact that 1 atm is equal to 1.013×10**5 N/m², the volume is 22.4 liters as shown in the calculation below

\[ V = \frac{n R_u T}{p} = \frac{(1\text{g-mol})(8.314\text{J/°K·g-mol})(273.16°\text{K})}{(1.013×10^5\text{N/m}^2)(\text{J/N·m})} = 0.0224\text{ m}^3 = 22.4\text{ L} \]  \hspace{3cm} (2)

**Example 2**

Determine the density of air at 20°C and 1 atm. Recalling that the molecular weight of air is 28.97 g/g-mole, the air density is easily calculated from the ideal gas law

\[ \rho = \frac{m}{V} = \frac{p M}{R_u T} = \frac{(1.013×10^5\text{N/m}^2)(28.97\text{g/g-mol})(\text{J/N·m})}{(8.314\text{J/°K·g-mol})(293.16°\text{K})(1000\text{g/kg})} = 1.204\text{ kg/m}^3 \]  \hspace{3cm} (3)

* I concede that the proper abbreviation of Kelvin is simply K, but I nonetheless choose to use to °K to avoid confusion with potassium (K) and, to a much lesser extent, kilo (k).
† For (combustion) fuel gases, STP is 60°F and 1 atm.