IDEAL GAS LAW

The ideal gas law is

$$p\,V = n\,R_u\,T = \frac{m}{M}\,R_u\,T = m\,R\,T = N\,k\,T \qquad (1)$$

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·lbf/°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/m^2, the volume is 22.4 liters as shown in the calculation below

$$V = \frac{n\,R_u\,T}{p} = \frac{(1\,\text{gmol})(8.314\,\text{J/°K}\cdot\text{gmol})(273.16\,\text{°K})}{(1.013\times10^5\ \text{N/m}^2)(\text{J/N}\cdot\text{m})} = 0.0224\ \text{m}^3 = 22.4\ \text{L} \qquad (2)$$

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

$$\rho = \frac{m}{V} = \frac{p\,M}{R_u\,T} = \frac{(1.013\times10^5\ \text{N/m}^2)(28.97\,\text{g/gmol})(\text{J/N}\cdot\text{m})}{(8.314\,\text{J/°K}\cdot\text{gmol})(293.16\,\text{°K})(1000\,\text{g/kg})} = 1.204\ \text{kg/m}^3 \qquad (3)$$

[*] 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.