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Chapter 1 • Introduction

1.1 A gas at 20°C may be *any* if it contains less than 10^{23} molecules per mm^3 . If Avogadro's number is 6.02×10^{23} molecules per mole, what air pressure does this represent?

Solution: The mass of one molecule of air may be computed as

$$m = \frac{\text{Molecular weight}}{\text{Avogadro's number}} = \frac{28.97 \text{ mol}^{-1}}{6.02 \times 10^{23} \text{ molecules/mol}} = 4.81 \times 10^{-23} \text{ g}$$

Then the density of air containing 10^{23} molecules per mm^3 is, in SI units,

$$\rho = \left(10^{23} \frac{\text{molecules}}{\text{mm}^3} \right) \left(4.81 \times 10^{-23} \frac{\text{g}}{\text{molecule}} \right) = 4.81 \times 10^{-11} \frac{\text{g}}{\text{mm}^3} = 4.81 \times 10^{-5} \frac{\text{kg}}{\text{m}^3}$$

Finally, from the perfect gas law, Eq. (1.13), at 20°C = 293 K, we obtain the pressure:

$$p = \rho RT = \left(4.81 \times 10^{-5} \frac{\text{kg}}{\text{m}^3} \right) \left(287 \frac{\text{m}^2}{\text{s}^2 \cdot \text{K}} \right) (293 \text{ K}) = 4.0 \text{ Pa} \quad \text{Ans.}$$

1.2 The earth's atmosphere can be modeled as a uniform layer of air of thickness 20 km and average density 1.0 kg/m^3 (see Table A-6). Use these values to estimate the total mass and total number of molecules of air in the entire atmosphere of the earth.

Solution: Let R_e be the earth's radius = 6377 km. Then the total mass of air in the atmosphere is

$$m_a = \int \rho \, dV = \rho_{\text{air}} (\text{Air Vol}) = \rho_{\text{air}} (4\pi R_e^2 h) \quad \text{Ans.}$$
$$= (1.0 \text{ kg/m}^3) (4\pi (6377 \text{ km})^2 (20 \text{ km})) = 4.1 \times 10^{18} \text{ kg}$$

Dividing by the mass of one molecule = 4.8×10^{-23} g (see Prob. 1.1 above), we obtain the total number of molecules in the earth's atmosphere:

$$N_{\text{molecules}} = \frac{m_{\text{atmosphere}}}{m_{\text{one molecule}}} = \frac{4.1 \times 10^{18} \text{ grams}}{4.8 \times 10^{-23} \text{ g/molecule}} = 1.3 \times 10^{44} \text{ molecules} \quad \text{Ans.}$$

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