Mole (unit for amount of a substance)

One **mole** (abbreviated mol) is defined as the amount of substance that contains as many atoms or molecules as there are in precisely 12 grams of carbon 12 (whose atomic mass is exactly 12 u). A simpler but equivalent definition is this: 1 mol is that number of grams of a substance numerically equal to the molecular mass (Section 13–1) of the substance. For example, the molecular mass of hydrogen gas (H_2) is 2.0 u (since each molecule contains two atoms of hydrogen and each atom has an atomic mass of 1.0 u). Thus 1 mol of H_2 has a mass of 2.0 g. Similarly, 1 mol of neon gas has a mass of 20 g, and 1 mol of CO_2 has a mass of $[12 + (2 \times 16)] = 44$ g since oxygen has atomic mass of 16 (see periodic Table inside the rear cover). The mole is the official unit of amount of substance in the SI system. In general, the number of moles, n, in a given sample of a pure substance is equal to the mass of the sample in grams divided by the molecular mass specified as grams per mole:

$$n\,(\text{mol}) = \frac{\text{mass (grams)}}{\text{molecular mass (g/mol)}}.$$

For example, the number of moles in 132 g of CO_2 (molecular mass 44 u) is

$$n = \frac{132\text{ g}}{44\text{ g/mol}} = 3.0\text{ mol}.$$

We can now write the proportion discussed above as an equation:

IDEAL GAS LAW

$$PV = nRT, \tag{13–3}$$

where n represents the number of moles and R is the constant of proportionality. R is called the **universal gas constant** because its value is found experimentally to be the same for all gases. The value of R, in several sets of units (only the first is the proper SI unit), is

Universal gas constant (in various units)

$$\begin{aligned} R &= 8.314\text{ J/(mol}\cdot\text{K)} \qquad \text{[SI units]}\\ &= 0.0821\text{ (L}\cdot\text{atm)/(mol}\cdot\text{K)}\\ &= 1.99\text{ calories/(mol}\cdot\text{K)}.^{\dagger} \end{aligned}$$

Equation 13–3 is called the **ideal gas law**, or the **equation of state for an ideal gas**. We use the term "ideal" because real gases do not follow Eq. 13–3 precisely, particularly at high pressure (and density) or when the gas is near the liquefaction point (= boiling point). However, at pressures less than an atmosphere or so, and when T is not close to the liquefaction point of the gas, Eq. 13–3 is quite accurate and useful for real gases.

CAUTION

Always give T in kelvins and P as absolute, not gauge, pressure

Always remember, when using the ideal gas law, that temperatures must be given in kelvins (K) and that the pressure P must always be *absolute* pressure, not gauge pressure (Section 10–4).

13–8 Problem Solving with the Ideal Gas Law

The ideal gas law is an extremely useful tool, and we now consider some Examples. We will often refer to "standard conditions" or "standard temperature and pressure" (STP), which means:

STP = 273 K, 1 atm

$$T = 273\text{ K}\ \ (0°\text{C})\ \ \text{and}\ \ P = 1.00\text{ atm} = 1.013 \times 10^5\text{ N/m}^2 = 101.3\text{ kPa}.$$

EXAMPLE 13–10 **Volume of one mol at STP.** Determine the volume of 1.00 mol of any gas, assuming it behaves like an ideal gas, at STP.

APPROACH We use the ideal gas law, solving for V.

SOLUTION We solve for V in Eq. 13–3:

$$V = \frac{nRT}{P} = \frac{(1.00\text{ mol})(8.314\text{ J/mol}\cdot\text{K})(273\text{ K})}{(1.013 \times 10^5\text{ N/m}^2)} = 22.4 \times 10^{-3}\text{ m}^3.$$

1 mol of gas at STP has $V = 22.4$ L.

Since 1 liter is 1000 cm^3 = 1×10^{-3} m^3, 1 mol of any gas has $V = 22.4$ L at STP.

[†]Calories will be defined in Section 14–1; sometimes it is useful to use R as given in terms of calories.