

$$\begin{aligned}
 M &= RT \left( \frac{\rho}{p} \right)_{p=0} = (0.082\,061\,5 \text{ L atm mol}^{-1} \text{ K}^{-1}) \times (273.15 \text{ K}) \times (1.42755 \text{ g L}^{-1} \text{ atm}^{-1}) \\
 &= \boxed{31.9987 \text{ g mol}^{-1}}
 \end{aligned}$$

The value obtained for  $R$  deviates from the accepted value by 0.005 per cent. The error results from the fact that only three data points are available and that a linear extrapolation was employed. The molar mass, however, agrees exactly with the accepted value, probably because of compensating plotting errors.

**E1.11(b)** The mass density  $\rho$  is related to the molar volume  $V_m$  by

$$V_m = \frac{M}{\rho}$$

where  $M$  is the molar mass. Putting this relation into the perfect gas law yields

$$pV_m = RT \quad \text{so} \quad \frac{pM}{\rho} = RT$$

Rearranging this result gives an expression for  $M$ ; once we know the molar mass, we can divide by the molar mass of phosphorus atoms to determine the number of atoms per gas molecule

$$M = \frac{RT\rho}{p} = \frac{(62.364 \text{ L Torr K}^{-1} \text{ mol}^{-1}) \times [(100 + 273) \text{ K}] \times (0.6388 \text{ g L}^{-1})}{120 \text{ Torr}} = 124 \text{ g mol}^{-1}$$

The number of atoms per molecule is

$$\frac{124 \text{ g mol}^{-1}}{31.0 \text{ g mol}^{-1}} = 4.00$$

suggesting a formula of  $\boxed{\text{P}_4}$

**E1.12(b)** Use the perfect gas equation to compute the amount; then convert to mass.

$$pV = nRT \quad \text{so} \quad n = \frac{pV}{RT}$$

We need the partial pressure of water, which is 53 per cent of the equilibrium vapour pressure at the given temperature and standard pressure.

$$\begin{aligned}
 p &= (0.53) \times (2.69 \times 10^3 \text{ Pa}) = 1.43 \times 10^3 \text{ Pa} \\
 \text{so } n &= \frac{(1.43 \times 10^3 \text{ Pa}) \times (250 \text{ m}^3)}{(8.3145 \text{ J K}^{-1} \text{ mol}^{-1}) \times (23 + 273) \text{ K}} = 1.43 \times 10^2 \text{ mol}
 \end{aligned}$$

$$\text{or } m = (1.43 \times 10^2 \text{ mol}) \times (18.0 \text{ g mol}^{-1}) = 2.61 \times 10^3 \text{ g} = \boxed{2.61 \text{ kg}}$$

**E1.13(b)** (a) The volume occupied by each gas is the same, since each completely fills the container. Thus solving for  $V$  from eqn 1.4 we have (assuming a perfect gas)

$$\begin{aligned}
 V &= \frac{n_J RT}{p_J} \quad n_{\text{Ne}} = \frac{0.225 \text{ g}}{20.18 \text{ g mol}^{-1}} \\
 &= 1.113 \times 10^{-2} \text{ mol}, \quad p_{\text{Ne}} = 66.5 \text{ Torr}, \quad T = 300 \text{ K}
 \end{aligned}$$

$$V = \frac{(1.113 \times 10^{-2} \text{ mol}) \times (62.36 \text{ L Torr K}^{-1} \text{ mol}^{-1}) \times (300 \text{ K})}{66.5 \text{ Torr}} = 3.137 \text{ L} = \boxed{3.14 \text{ L}}$$

# Physical Chemistry Atkins 7th Edition Solutions Manual

**R Sandford**



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