

1. The density of water at 20 °C is 0.998 g·cm<sup>-3</sup>. What height would the column of liquid be in a water barometer when the atmospheric pressure corresponds to 760 mm of mercury?

$$760. \text{ mm} = 1.01 \times 10^5 \text{ Pa} = 1.01 \times 10^5 \text{ kg} \cdot \text{m}^{-1} \cdot \text{s}^{-2}$$

$$P = dhg$$

$$h = \frac{P}{dg} = \frac{1.01 \times 10^5 \text{ kg} \cdot \text{m}^{-1} \cdot \text{s}^{-2}}{998 \text{ kg} \cdot \text{m}^{-3} \times 9.80665 \text{ m} \cdot \text{s}^{-2}} = 10.3 \text{ m}$$

Note: This also corresponds to the theoretical maximum suction lift of a pump at that pressure.

2. In a petroleum refinery, a 750. L container containing ethylene gas at 1.00 bar was compressed isothermally to 5.00 bar. What was the final volume of the container?

Since only volume and pressure are changing, use Boyle's Law:

$$p_2V_2 = P_1V_1$$

$$V_2 = \frac{P_1V_1}{P_2} \cong \frac{1.00 \text{ bar} \times 750. \text{ L}}{5.00 \text{ bar}} = 150. \text{ L}$$

3. A helium weather balloon was filled at -20. °C and at a certain pressure to a volume of 2.5 × 10<sup>4</sup> L with 1.2 × 10<sup>3</sup> mol He. What is the molar volume of helium under those conditions?

The units are molar volume or L·mol<sup>-1</sup>, so use Avogadro's principle:

$$V_m = \frac{V}{n}$$

$$V_m = \frac{2.5 \times 10^4 \text{ L}}{1.2 \times 10^3 \text{ mol}} = 20.83 \text{ L} \cdot \text{mol}^{-1}$$

4. Air is a source of reactants for many chemical processes. To determine the amount of air required for these reactions, it is helpful to know the partial pressures of the components. A particular sample of dry air of total mass 1.00 g consists almost entirely of 0.760 g of nitrogen and 0.240 g of oxygen. Calculate the partial pressures of these gases when the total pressure is 0.87 atm.

For finding the partial pressure of a gas in a mixture of gases, use  $P_A = x_A P$ .

To find the mole fraction of each gas, first find the moles of each gas from the given mass.

$$\frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} \times \frac{0.760 \text{ g N}_2}{1} = 0.0271 \text{ mol N}_2$$

$$\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{0.240 \text{ g O}_2}{1} = 0.00750 \text{ mol O}_2$$

$$0.0271 \text{ mol N}_2 + 0.00750 \text{ mol O}_2 = 0.0346 \text{ mol total}$$

Find the mole fraction of each gas.

$$x_{\text{N}_2} = \frac{n_{\text{N}_2}}{n_{\text{N}_2} + n_{\text{O}_2}} = \frac{0.0271 \text{ mol N}_2}{0.0346 \text{ mol}} = 0.7830$$

$$x_{\text{O}_2} = \frac{n_{\text{O}_2}}{n_{\text{N}_2} + n_{\text{O}_2}} = \frac{0.00750 \text{ mol O}_2}{0.0346 \text{ mol}} = 0.217$$

Find the partial pressures of the gases using the mole fractions and the total pressure (0.87 atm).

$$P_{\text{N}_2} = x_{\text{N}_2}P = 0.783 \times 0.87 \text{ atm} = \mathbf{0.68 \text{ atm N}_2}$$

$$P_{\text{O}_2} = x_{\text{O}_2}P = 0.217 \times 0.87 \text{ atm} = \mathbf{0.19 \text{ atm O}_2}$$

- 5. Divers exploring a shipwreck and wishing to avoid the narcosis associated with breathing nitrogen under high pressure switch to a neon-oxygen gas mixture containing 141.2 g of oxygen and 335.0 g of neon. The pressure in the gas tanks is 50.0 atm. What is the partial pressure of oxygen in the tanks?**

To find the partial pressure of a gas in a mixture of gases use  $P_A = x_A P$ .

In order to find the mole fraction of each gas, first find the moles of each gas from the given mass.

$$\frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}} \times \frac{335.0 \text{ g Ne}}{1} = 16.60 \text{ mol Ne}$$

$$\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{141.2 \text{ g O}_2}{1} = 4.413 \text{ mol O}_2$$

$$16.60 \text{ mol Ne} + 4.413 \text{ mol O}_2 = 21.01 \text{ mol total}$$

$$x_{\text{Ne}} = \frac{n_{\text{Ne}}}{n_{\text{Ne}} + n_{\text{O}_2}} = \frac{16.60 \text{ mol Ne}}{21.01 \text{ mol}} = 0.7900$$

$$x_{\text{O}_2} = \frac{n_{\text{O}_2}}{n_{\text{Ne}} + n_{\text{O}_2}} = \frac{4.413 \text{ mol O}_2}{21.01 \text{ mol}} = 0.2100$$

Find the partial pressures of the gases using the mole fractions and the total pressure (50.0 atm).

$$P_{\text{Ne}} = x_{\text{Ne}}P = 0.7900 \times 50.0 \text{ atm} = 39.5 \text{ atm Ne}$$

$$P_{\text{O}_2} = x_{\text{O}_2}P = 0.2100 \times 50.0 \text{ atm} = \mathbf{10.5 \text{ atm O}_2}$$

6. A weather balloon is filled with helium gas at 20. °C and 1.00 atm. The volume of the balloon is 250. L. When the balloon rises to a layer of air where the temperature is -30. °C, it has expanded to 800. L. What is the pressure of the atmosphere at that point?

This problem has three variables that change: pressure, volume, and temperature. Therefore, we can use the combined gas law.

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

A table of variables and values:

$$P_2 = ?$$

$$P_1 = 1.00 \text{ atm}$$

$$V_2 = 800. \text{ L}$$

$$V_1 = 250. \text{ L}$$

$$T_2 = 273.15 + -30. \text{ °C} = 243 \text{ K}$$

$$T_1 = 273.15 + 20. \text{ °C} = 293$$

$$\text{Solve for } P_2 \cdot P_2 = \frac{P_1 V_1 T_2}{T_1 V_2} \cong \frac{1.00 \text{ atm} \times 250. \text{ L} \times 243 \text{ K}}{293 \text{ K} \times 800. \text{ L}} = \mathbf{0.259 \text{ atm}}$$

7. Calculate the volume occupied by 2.0 g of helium at 25 °C and 1.0 atm.

Rearrange the ideal gas law to solve for volume:

$$V = \frac{nRT}{P}$$

$$R = 8.314 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1} \text{ or } 0.08206 \text{ atm} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$$

Change the temperature:  $273.15 + 25. \text{ °C} = 298 \text{ K}$

Find moles of He:

$$\frac{1 \text{ moles He}}{4.00 \text{ g He}} \times \frac{2.0 \text{ g He}}{1} = 0.50 \text{ mol He}$$

Substitute data into the formula:

$$V \cong \frac{0.50 \text{ mol} \times 0.08206 \text{ atm} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \times 298 \text{ K}}{1.0 \text{ atm}} = \mathbf{12 \text{ L}}$$

- 8. The volatile organic compound geraniol, a component of oil of roses, is used in perfumery. The density of the vapor at 260. °C and 103 Torr is 0.480 g·L<sup>-1</sup>. What is the molar mass of geraniol?**

When working with densities of gases use:

$$d = \frac{MP}{RT}$$

Since we're solving for M we'll need R to match our pressure, so we use 62.364 Torr·K<sup>-1</sup>·mol<sup>-1</sup> and change the temperature to K, 273.15 + 260. °C = 533 K.

Solve for molar mass using the equation:

$$M = \frac{dRT}{P}$$

Insert the values into the equation:

$$\frac{0.480 \text{ g} \cdot \text{L}^{-1} \times 62.364 \text{ Torr} \cdot \text{K}^{-1} \cdot \text{mol}^{-1} \times 533 \text{ K}}{103 \text{ Torr}} = 155 \text{ g} \times \text{mol}^{-1}$$

- 9. The Codex Ebers, an Egyptian medical papyrus, describes the use of garlic as an antiseptic. Chemists today have verified that the oxide of diallyl disulfide (the volatile compound responsible for garlic odor) is a powerful antibacterial agent. At 177 °C and 200. Torr, a sample of diallyl disulfide vapor has a density of 1.04 g·L<sup>-1</sup>. What is the molar mass of diallyl disulfide?**

When working with densities of gases use:

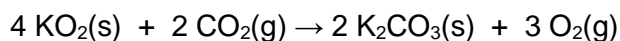
$$d = \frac{MP}{RT}$$

Since we're solving for M we'll need R to match our pressure, so we use 62.364 Torr·K<sup>-1</sup>·mol<sup>-1</sup> and change the temperature to K, 273.15 + 177. °C = 450. K  
Solve for the molar mass using the equation:

$$M = \frac{dRT}{P}$$

$$\frac{1.04 \text{ g} \cdot \text{L}^{-1} \times 62.364 \text{ Torr} \cdot \text{K}^{-1} \cdot \text{mol}^{-1} \times 450. \text{ K}}{200. \text{ Torr}} = 146 \text{ g} \cdot \text{L}^{-1}$$

- 10. The carbon dioxide generated by the personnel in the artificial atmosphere of submarines and spacecraft must be removed from the air and the oxygen recovered. Submarine design teams have investigated the use of potassium superoxide, KO<sub>2</sub>, as an air purifier because this compound reacts with carbon dioxide and releases oxygen. Calculate the mass of KO<sub>2</sub> needed to react with 50. L of carbon dioxide at 25 °C and 1.0 atm.**



The strategy: moles CO<sub>2</sub> → moles KO<sub>2</sub> → g KO<sub>2</sub>

Information given for  $\text{CO}_2$  tells us we are at SATP so we look up molar volume to find,  $V_m = 24.47 \text{ L} \cdot \text{mol}^{-1}$ .

To find moles of  $\text{CO}_2$  we use  $V_m = V/n$  rearranged to  $n = V/V_m$ . Plug in  $V$  and  $V_m$  to get  $n = 50. \text{ L CO}_2 / 24.47 \text{ L} \cdot \text{mol}^{-1} \text{ CO}_2$ .

The balanced formula provides mole  $\text{KO}_2$  to mole  $\text{CO}_2$  so we can use the mole ratio:  
 $4 \text{ mol KO}_2 = 2 \text{ mol CO}_2$

Use molar mass to find grams  $\text{KO}_2$ ,  $1 \text{ mol KO}_2 = 71.1 \text{ g KO}_2$

$$\frac{4 \text{ mol KO}_2}{2 \text{ mol CO}_2} \times \frac{50. \text{ L CO}_2}{24.47 \text{ L} \cdot \text{mol}^{-1} \text{ CO}_2} \times \frac{71.1 \text{ g KO}_2}{1 \text{ mol KO}_2} = \mathbf{290 \text{ g KO}_2}$$

### 11. Estimate the root mean square speed of $\text{CH}_4$ molecules at $25^\circ\text{C}$ .

The temperature is  $298 \text{ K}$  and the molar mass of  $\text{CH}_4$  is  $16.04 \text{ g} \cdot \text{mol}^{-1}$  (corresponding to  $1.604 \times 10^{-2} \text{ kg} \cdot \text{mol}^{-1}$ ) and  $R = 8.314 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$ .

$$v_{\text{rms}} = \left( \frac{3 \times 8.314 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1} \times 298 \text{ K}}{1.604 \times 10^{-2} \text{ kg} \cdot \text{mol}^{-1}} \right)^{\frac{1}{2}} = \mathbf{681 \text{ m} \cdot \text{s}^{-1}}$$

### 12. A cylinder of volume $2.00 \text{ L}$ contains $1.00 \text{ mol He(g)}$ at $30^\circ\text{C}$ .

Which process does more work on the surroundings: allowing the gas to expand isothermally to  $4.00 \text{ L}$  against a constant external pressure of  $1.00 \text{ atm}$ , or allowing it to expand reversibly and isothermally to the same final volume?

a) Irreversible path:  $w = -P_{\text{ex}}\Delta V$

$$w = -(1.00 \text{ atm}) \times (2.00 \text{ L}) = -1.00 \times 2.00 \text{ L} \cdot \text{atm} \times \frac{101.325 \text{ J}}{1 \text{ L} \cdot \text{atm}} = -203 \text{ J}$$

b) Reversible isothermal path:

$$w = -nRT \ln \frac{V_2}{V_1}$$

$$w = (-1.00 \text{ mol}) \times (8.3145 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}) \times (303 \text{ K}) \times \ln \frac{4.00 \text{ L}}{2.00 \text{ L}} = -1750 \text{ J}$$

**The gas does more work in the reversible process.**