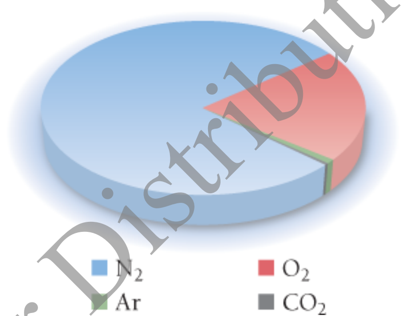
10.7: Mixtures of Gases and Partial Pressures

Key Concept Video **Mixtures of Gases and Partial Pressures**

Many gas samples are not pure; they are mixtures of gases. Dry air, for example, is a mixture containing nitrogen, oxygen, argon, carbon dioxide, and a few other gases in trace amounts (Table 10.2).

Table 10.2 **Composition of Dry Air**

Gas	Percent by Volume (%)
Nitrogen (N_2)	78
Oxygen (O_2)	21
Argon (Ar)	0.9
Carbon dioxide (CO_2)	0.04



According to kinetic molecular theory, the particles in a gas mixture have negligible size and they do not interact. Consequently, each of the components in an ideal gas mixture acts independently of the others. For example, the nitrogen molecules in air exert a certain pressure—78% of the total pressure—that is independent of the other gases in the mixture. Likewise, the oxygen molecules in air exert a certain pressure—21% of the total pressure—that is also independent of the other gases in the mixture. The pressure due to any individual component in a gas mixture is its **partial pressure (P_n)**. We can calculate partial pressure from the ideal gas law by assuming that each gas component acts independently:

$$P_n = n_n \frac{RT}{V}$$

For a multicomponent gas mixture, we calculate the partial pressure of each component from the ideal gas law and the number of moles of that component (n_n) as follows:

[10.16]

$$P_a = n_a \frac{\mathrm{RT}}{\mathrm{V}}; \quad P_b = n_b \frac{RT}{V}; \quad P_c = n_c \frac{RT}{V}; \ldots$$

According to kinetic molecular theory, the only property that distinguishes one type of particle from another is its mass. However, even particles of different masses have the same average kinetic energy at a given temperature, so they exert the same force upon collision with a surface. Consequently, adding different kinds of gases to a gas mixture has the same effect on pressure as simply adding more particles. The partial pressures of all the components sum to the overall pressure:

[10.17]

$$P_{\text{total}} = P_{\text{a}} + P_{\text{b}} + P_{\text{c}} + \ldots$$

where P_{total} is the total pressure and P_{a}, P_{b}, P_{c}, … , are the partial pressures of the components. This relationship is known as **Dalton's law of partial pressures**.

Combining Equations 10.16 and 10.17, we get:

[10.18]

$$\begin{aligned} P_{\text{total}} &= P_{\text{a}} + P_{\text{b}} + P_{\text{c}} + \ldots \\ &= n_{\text{a}}\frac{RT}{V} + n_{\text{b}}\frac{RT}{V} + n_{\text{c}}\frac{RT}{V} + \ldots \\ &= (n_{\text{a}} + n_{\text{b}} + n_{\text{c}} + \ldots)\frac{RT}{V} \\ &= (n_{\text{total}})\frac{RT}{V} \end{aligned}$$

The total number of moles in the mixture, when substituted into the ideal gas law, indicates the total pressure of the sample.

If we divide Equation 10.16 by Equation 10.18, we get:

[10.19]

$$\frac{P_{\text{a}}}{P_{\text{total}}} = \frac{n_{\text{a}}\,\cancel{(RT/V)}}{n_{\text{total}}\,\cancel{(RT/V)}} = \frac{n_{\text{a}}}{n_{\text{total}}}$$

The quantity $n_{\text{a}}/n_{\text{total}}$, the number of moles of a component in a mixture divided by the total number of moles in the mixture, is the **mole fraction** (χ_a):

[10.20]

$$\chi_a = \frac{n_{\text{a}}}{n_{\text{total}}}$$

Rearranging Equation 10.19 and substituting the definition of mole fraction (Equation 10.20) gives us:

$$\begin{aligned} \frac{P_{\text{a}}}{P_{\text{total}}} &= \frac{n_{\text{a}}}{n_{\text{total}}} \\ P_{\text{a}} &= \frac{n_{\text{a}}}{n_{\text{total}}}P_{\text{total}} = \chi_{\text{a}}P_{\text{total}} \end{aligned}$$

or more simply:

[10.21]

$$P_{\text{a}} = \chi_{\text{a}}P_{\text{total}}$$

The partial pressure of a component in a gaseous mixture is its mole fraction multiplied by the total pressure. For gases, the mole fraction of a component is equivalent to its percent by volume divided by 100%. Therefore, based on Table 10.2, we calculate the partial pressure of nitrogen (P_{N_2}) in air at 1.00 atm to be:

$$\begin{aligned} P_{\text{N}_2} &= 0.78 \times 1.00\ \text{atm} \\ &= 0.78\ \text{atm} \end{aligned}$$

Likewise, the partial pressure of oxygen in air at 1.00 atm is 0.21 atm, and the partial pressure of argon in air is 0.01 atm. Applying Dalton's law of partial pressures to air at 1.00 atm:

$$\begin{aligned} P_{\text{total}} &= P_{N_2} + P_{O_2} + P_{Ar} \\ P_{\text{total}} &= 0.78 \text{ atm} + 0.21 \text{ atm} + 0.01 \text{ atm} \\ &= 1.00 \text{ atm} \end{aligned}$$

For these purposes, we ignore the contribution of CO_2 and other trace gases in air because they are so small.

Conceptual Connection 10.5 **Partial Pressures**

Example 10.9 Total Pressure and Partial Pressures

A 1.00-L mixture of helium, neon, and argon has a total pressure of 662 mmHg at 298 K. If the partial pressure of helium is 341 mmHg and the partial pressure of neon is 112 mmHg, what mass of argon is present in the mixture?

SORT The problem gives you partial pressures for two of the three components in a gas mixture, along with the total pressure, the volume, and the temperature, and asks you to find the mass of the third component.

GIVEN:

$$\begin{aligned} P_{\text{He}} &= 341 \text{ mmHg}, \\ P_{\text{Ne}} &= 112 \text{ mmHg}, \\ P_{\text{total}} &= 662 \text{ mmHg}, \\ V &= 1.00 \text{ L}, T = 298 \text{ K} \end{aligned}$$

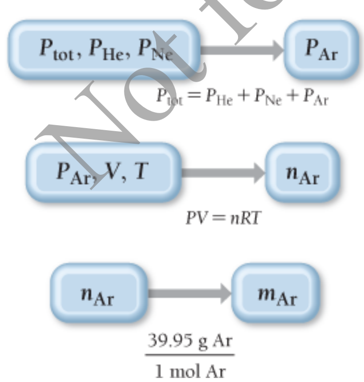
FIND: m_{Ar}

STRATEGIZE You can find the mass of argon from the number of moles of argon, which you can calculate from the partial pressure of argon and the ideal gas law. Begin by using Dalton's law to determine the partial pressure of argon.

Then use the partial pressure of argon together with the volume of the sample and the temperature to find the number of moles of argon.

Finally, use the molar mass of argon to calculate the mass of argon from the number of moles of argon.

CONCEPTUAL PLAN



RELATIONSHIPS USED

$$\begin{aligned} P_{\text{total}} &= P_{\text{He}} + P_{\text{Ne}} + P_{\text{Ar}} \text{ (Dalton's law)} \\ PV &= nRT \text{ (ideal gas law)} \\ \text{molar mass Ar} &= 39.95 \text{ g/mol} \end{aligned}$$

SOLVE Follow the conceptual plan. To find the partial pressure of argon, solve the equation for P_{Ar} and substitute the values of the other partial pressures to calculate P_{Ar}.

Convert the partial pressure from mmHg to atm and use it in the ideal gas law to calculate the amount of argon in moles.

Use the molar mass of argon to convert from amount of argon in moles to mass of argon.

SOLUTION

$$\begin{aligned} P_{\text{total}} &= P_{\text{He}} + P_{\text{Ne}} + P_{\text{Ar}} \\ P_{\text{Ar}} &= P_{\text{total}} - P_{\text{He}} - P_{\text{Ne}} \\ &= 662 \text{ mmHg} - 341 \text{ mmHg} - 112 \text{ mmHg} \\ &= 209 \text{ mmHg} \end{aligned}$$

$$209 \ \cancel{\text{mmHg}} \times \frac{1 \text{ atm}}{760 \ \cancel{\text{mmHg}}} = 0.275 \text{ atm}$$

$$n = \frac{PV}{RT} = \frac{0.275 \ \cancel{\text{atm}} \left(1.00 \ \cancel{\text{L}}\right)}{0.08206 \dfrac{\cancel{\text{L}} \cdot \cancel{\text{atm}}}{\text{mol} \cdot \cancel{\text{K}}} \left(298 \ \cancel{\text{K}}\right)} = 1.1\underline{2}5 \times 10^{-2} \text{ mol Ar}$$

$$1.1\underline{2}5 \times 10^{-2} \ \cancel{\text{mol Ar}} \times \frac{39.95 \text{ g Ar}}{1 \ \cancel{\text{mol Ar}}} = 0.449 \text{ g Ar}$$

CHECK The units of the answer are correct. The magnitude of the answer makes sense because the volume is 1.0 L, which at STP would contain about 1/22 mol. Because the partial pressure of argon in the mixture is about 1/3 of the total pressure, you can roughly estimate about 1/66 of one molar mass of argon, which is fairly close to your answer.

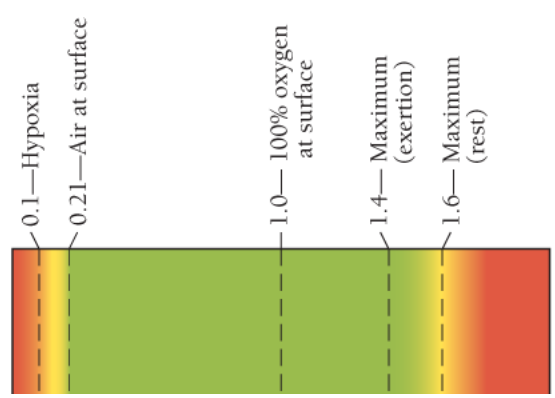
FOR PRACTICE 10.9 A sample of hydrogen gas is mixed with water vapor. The mixture has a total pressure of 755 torr, and the water vapor has a partial pressure of 24 torr. What amount (in moles) of hydrogen gas is contained in 1.55 L of this mixture at 298 K?

Deep-Sea Diving and Partial Pressures

Our lungs have evolved to breathe oxygen at a partial pressure of $P_{O_2} = 0.21$ atm. If the total air pressure decreases—when a person climbs a mountain, for example—the partial pressure of oxygen also decreases. On top of Mount Everest, where the total pressure is 0.311 atm, the partial pressure of oxygen is only 0.065 atm. Low oxygen levels produce a physiological condition called **hypoxia** or oxygen starvation (Figure 10.17). Mild hypoxia causes dizziness, headache, and shortness of breath. Severe hypoxia, which occurs when (P_{O_2}) drops below 0.1 atm, may result in unconsciousness or even death. For this reason, climbers hoping to make the summit of Mount Everest usually carry oxygen to breathe.

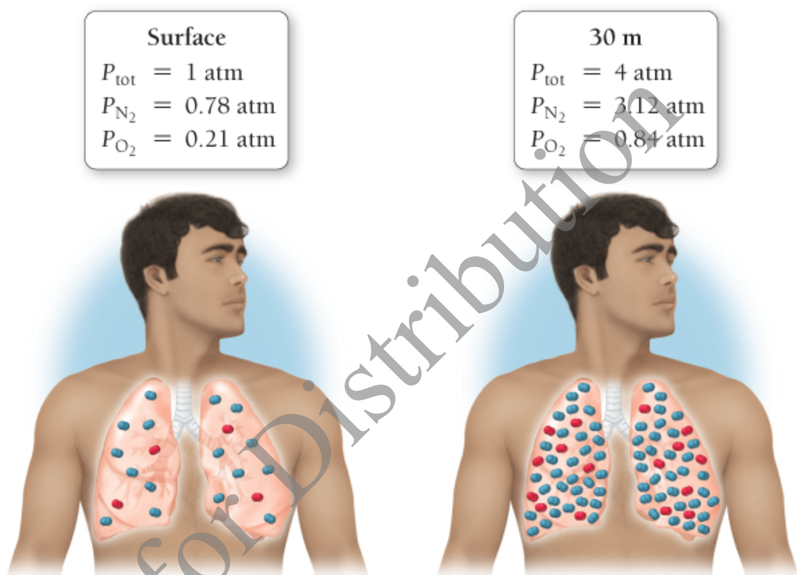
Figure 10.17 Oxygen Partial Pressure Limits

The partial pressure of oxygen in air at sea level is 0.21 atm. Partial pressures of oxygen below 0.1 atm and above 1.4 atm are dangerous to humans.



While not as dangerous as a lack of oxygen, too much oxygen can also cause physiological problems. Recall from Section 10.4 that scuba divers breathe pressurized air. At 30 m, a scuba diver breathes air at a total pressure of 4.0 atm, which means P_{O_2} is about 0.84 atm. This elevated partial pressure of oxygen raises the density of oxygen molecules in the lungs, resulting in a higher concentration of oxygen in body tissues. When P_{O_2} increases beyond 1.4 atm, the increased oxygen concentration in body tissues causes a condition called **oxygen toxicity** that results in muscle twitching, tunnel vision, and convulsions. Divers who venture too deep without proper precautions have drowned because of oxygen toxicity.

A second problem associated with breathing pressurized air is the increase of nitrogen in the lungs. At 30 m, a scuba diver breathes nitrogen at $P_{N_2} = 3.12$ atm, which increases the nitrogen concentration in body tissues and fluids. When P_{N_2} increases beyond about 4 atm, a condition called **nitrogen narcosis** or *rapture of the deep* results. Divers describe the effects of this condition as similar to being inebriated or drunk. A diver breathing compressed air at 60 m feels as if he has consumed too much wine.



When a diver breathes compressed air, the abnormally high partial pressure of oxygen in the lungs leads to an elevated concentration of oxygen in body tissues.

To avoid oxygen toxicity and nitrogen narcosis, deep-sea divers—those who descend beyond 50 m—breathe specialized mixtures of gases. One common mixture is heliox, a mixture of helium and oxygen. These mixtures usually contain a smaller percentage of oxygen than is typically found in air, thereby lowering the risk of oxygen toxicity. Heliox also contains helium instead of nitrogen, eliminating the risk of nitrogen narcosis.

Example 10.10 Partial Pressures and Mole Fractions

A 12.5-L scuba diving tank contains a helium-oxygen (heliox) mixture of 24.2 g of He and 4.32 g of O_2 at 298 K. Calculate the mole fraction and partial pressure of each component in the mixture and the total pressure of the mixture.

SORT The problem gives the masses of two gases in a mixture and the volume and temperature of the mixture. You are to find the mole fraction and partial pressure of each component, as well as the total pressure.

GIVEN: $m_{\mathrm{He}} = 24.2\ g, m_{\mathrm{O_2}} = 4.32$ g,
$V = 12.5$ L, $T = 298$ K

FIND: $\chi_{\mathrm{He}}, \chi_{\mathrm{O_2}}, P_{\mathrm{He}}, P_{\mathrm{O_2}}, P_{\mathrm{total}}$

STRATEGIZE The conceptual plan has several parts. To calculate the mole fraction of each component, you must first find the number of moles of each component. In the first part of the conceptual plan, convert the masses to moles using the molar masses.

In the second part, calculate the mole fraction of each component using the mole fraction definition (Equation 10.20).

To calculate *partial pressures* calculate the *total pressure* and then use the mole fractions from the previous calculation to determine the partial pressures. Calculate the total pressure from the sum of the moles of both components. (Alternatively, you can calculate the partial pressures of the components individually, using the number of moles of each component and adding them to obtain the total pressure.)

Use the mole fractions of each component and the total pressure to calculate the partial pressure of each component.

CONCEPTUAL PLAN

$$m_{\mathrm{He}} \xrightarrow[\frac{1\ \mathrm{mol\ He}}{4.00\ \mathrm{g\ He}}]{} n_{\mathrm{He}} \qquad m_{\mathrm{O_2}} \xrightarrow[\frac{1\ \mathrm{mol\ O_2}}{32.00\ \mathrm{g\ O_2}}]{} n_{\mathrm{O_2}}$$

$$\begin{aligned}
\chi_{\mathrm{He}} &= \frac{n_{\mathrm{He}}}{n_{\mathrm{He}} + n_{\mathrm{O_2}}}; \chi_{\mathrm{O_2}} = \frac{n_{\mathrm{O_2}}}{n_{\mathrm{He}} + n_{\mathrm{O_2}}} \\
P_{\mathrm{total}} &= \frac{(n_{\mathrm{He}} + n_{\mathrm{O_2}})RT}{V} \\
P_{\mathrm{He}} &= \chi_{\mathrm{He}} P_{\mathrm{total}}; \quad P_{\mathrm{O_2}} = \chi_{\mathrm{O_2}} P_{\mathrm{total}}
\end{aligned}$$

RELATIONSHIPS USED

$$\begin{aligned}
\chi_{\mathrm{a}} &= n_{\mathrm{a}}/n_{\mathrm{total}} \text{ (mole fraction definition)} \\
P_{\mathrm{total}} V &= n_{\mathrm{total}} RT \text{ (ideal gas law)} \\
P_{\mathrm{a}} &= \chi_{\mathrm{a}} P_{\mathrm{total}}
\end{aligned}$$

SOLVE Follow the plan to solve the problem. Begin by converting each of the masses to amounts in moles.

Calculate each of the mole fractions.

Calculate the total pressure.

Finally, calculate the partial pressure of each component.

SOLUTION

$$\begin{aligned}
24.2\ \cancel{\mathrm{g\ He}} \times \frac{1\ \mathrm{mol\ He}}{4.00\ \cancel{\mathrm{g\ He}}} &= 6.05\ \mathrm{mol\ He} \\
4.32\ \cancel{\mathrm{g\ O_2}} \times \frac{1\ \mathrm{mol\ O_2}}{32.00\ \cancel{\mathrm{g\ O_2}}} &= 0.135\ \mathrm{mol\ O_2}
\end{aligned}$$

$$\begin{aligned}
\chi_{\mathrm{He}} &= \frac{n_{\mathrm{He}}}{n_{\mathrm{He}} + n_{\mathrm{O_2}}} = \frac{6.05}{6.06 + 0.135} = 0.97\underline{8}17 \\
\chi_{\mathrm{O_2}} &= \frac{n_{\mathrm{O_2}}}{n_{\mathrm{He}} + n_{\mathrm{O_2}}} = \frac{0.135}{6.05 + 0.135} = 0.021\underline{8}27 \\
P_{\mathrm{total}} &= \frac{(n_{\mathrm{He}} + n_{\mathrm{O_2}})RT}{V} \\
&= \frac{\left(6.05\ \cancel{\mathrm{mol}} + 0.135\ \cancel{\mathrm{mol}}\right)\left(0.08206 \frac{\cancel{\mathrm{L}} \cdot \mathrm{atm}}{\cancel{\mathrm{mol}} \cdot \cancel{\mathrm{K}}}\right)\left(298\ \cancel{\mathrm{K}}\right)}{12.5\ \mathrm{L}} \\
&= 12.\underline{0}99\ \mathrm{atm}
\end{aligned}$$

$$P_{He} = \chi_{He}P_{total} = 0.97\underline{8}17 \times 12.\underline{0}99 \text{ atm}$$
$$= 11.8 \text{ atm}$$
$$P_{O_2} = \chi_{O_2}P_{total} = 0.021\underline{8}27 \times 12.\underline{0}99 \text{ atm}$$
$$= 0.264 \text{ atm}$$

CHECK The units of the answers are correct, and the magnitudes are reasonable.

FOR PRACTICE 10.10 A diver breathes a heliox mixture with an oxygen mole fraction of 0.050. What must the total pressure be for the partial pressure of oxygen to be 0.21 atm?

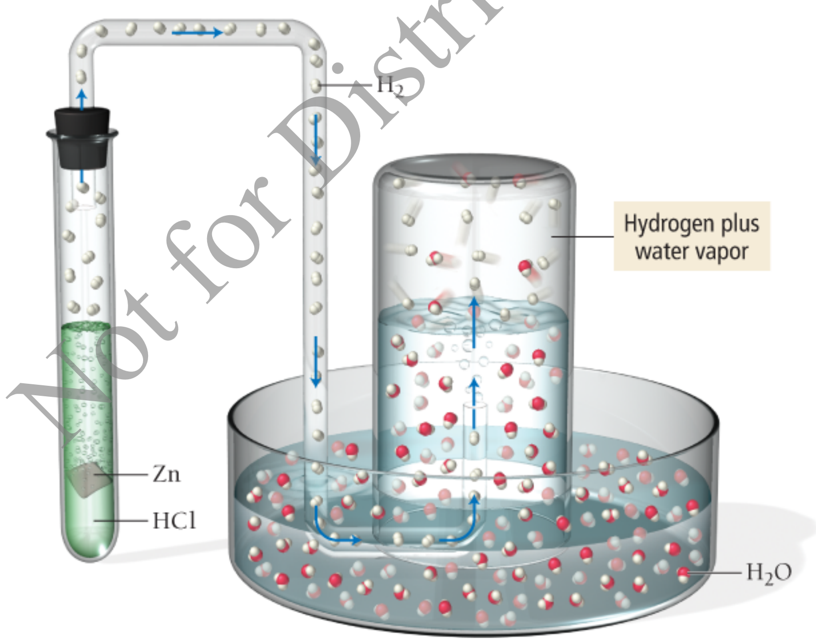
Collecting Gases over Water

When the desired product of a chemical reaction is a gas, we can collect the gas by the displacement of water. For example, suppose we use the reaction of zinc with hydrochloric acid as a source of hydrogen gas:

To collect the gas, we can set up an apparatus like the one shown in Figure 10.18. As the hydrogen gas forms, it bubbles through the water and gathers in the collection flask. The hydrogen gas collected in this way is not pure. It is mixed with water vapor because some water molecules evaporate and mix with the hydrogen molecules.

Figure 10.18 Collecting a Gas over Water

When we collect the gaseous product of a chemical reaction over water, the product molecules (in this case H_2) are mixed with water molecules. The pressure of those water molecules is equal to the vapor pressure of water at that temperature. The partial pressure of the product is the total pressure minus the partial pressure of water.



The partial pressure of water in the mixture, which we call its **vapor pressure**, depends on temperature (Table 10.3). Vapor pressure increases with increasing temperature because higher temperatures cause more water molecules to evaporate.

$$\text{Zn}(s) + 2\text{ HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)$$

We will discuss vapor pressure in detail in Chapter 11.

Appendix II includes a more complete table of the vapor pressure of water versus temperature.

Table 10.3 Vapor Pressure of Water versus Temperature

Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)
0	4.58	55	118.2
5	6.54	60	149.6
10	9.21	65	187.5
15	12.79	70	233.7
20	17.55	75	289.1
25	23.78	80	355.1
30	31.86	85	433.6
35	42.23	90	525.8
40	55.40	95	633.9
45	71.97	100	760.0
50	92.6		

Suppose we collect the hydrogen gas over water at a total pressure of 758.2 mmHg and a temperature of 25 °C. What is the partial pressure of the hydrogen gas? We know that the total pressure is 758.2 mmHg and that the partial pressure of water is 23.78 mmHg (its vapor pressure at 25 °C):

$$\begin{aligned} P_{\text{total}} &= P_{H_2} + P_{H_2O} \\ 758.2\text{ mmHg} &= P_{H_2} + 23.78\text{ mmHg} \end{aligned}$$

Therefore:

$$\begin{aligned} P_{H_2} &= 758.2\text{ mmHg} - 23.78\text{ mmHg} \\ &= 734.4\text{ mmHg} \end{aligned}$$

The partial pressure of the hydrogen in the mixture is 734.4 mmHg.

Example 10.11 Collecting Gases over Water

In order to determine the rate of photosynthesis, the oxygen gas emitted by an aquatic plant is collected over water at a temperature of 293 K and a total pressure of 755.2 mmHg. Over a specific time period, a total of 1.02 L of gas is collected. What mass of oxygen gas (in grams) forms?

SORT The problem gives the volume of gas collected over water as well as the temperature and the pressure. You are asked to find the mass in grams of oxygen that forms.

GIVEN:
$V = 1.02\text{ L}$, $P_{\text{total}} = 755.2\text{ mmHg}$,
$T = 293\text{K}$

FIND: g O_2

STRATEGIZE You can find the mass of oxygen from moles of oxygen, which you can get from the ideal gas law and the partial pressure of oxygen. Find the partial pressure of oxygen by subtracting the partial pressure of water at 293 K (20 °C) from the total pressure.

Use the ideal gas law to determine the number of moles of oxygen from its partial pressure, volume, and temperature.

temperature.

Then use the molar mass of oxygen to convert the number of moles to grams.

CONCEPTUAL PLAN

$$P_{O_2} = P_{total} - P_{H_2O}\ (20\ ^\circ C)$$

$$P_{O_2}, V, T \xrightarrow[P_{O_2}V = n_{O_2}RT]{} n_{O_2}$$

$$n_{O_2} \xrightarrow[\frac{32.00\ g\ O_2}{mol\ O_2}]{} g\ O_2$$

RELATIONSHIPS USED

$$\begin{aligned} P_{total} &= P_a + P_b + P_c + \ldots\ \text{(Dalton's law)} \\ PV &= nRT\ \text{(ideal gas law)} \end{aligned}$$

SOLVE Follow the conceptual plan to solve the problem. Begin by calculating the partial pressure of oxygen in the oxygen/water mixture. You can find the partial pressure of water at 20 °C in Table 10.3.

Next, solve the ideal gas law for number of moles.

Before substituting into the ideal gas law, convert the partial pressure of oxygen from mmHg to atm.

Substitute into the ideal gas law to find the number of moles of oxygen.

Finally, use the molar mass of oxygen to convert to grams of oxygen.

SOLUTION

$$\begin{aligned} P_{O_2} &= P_{total} - P_{H_2O}\ (20\ ^\circ C) \\ &= 755.2\ mmHg - 17.55\ mmHg = 737.\underline{6}5\ mmHg \\ n_{O_2} &= \frac{P_{O_2}V}{RT} \end{aligned}$$

$$737.65\ \cancel{mmHg} \times \frac{1\ atm}{760\ \cancel{mmHg}} = 0.970\underline{5}9\ atm$$

$$n_{O_2} = \frac{P_{O_2}V}{RT} = \frac{0.970\underline{5}9\ \cancel{atm}\ (1.02\ \cancel{L})}{0.08206\ \frac{\cancel{L}\cdot\cancel{atm}}{mol\cdot\cancel{K}}(293\ K)} = 4.1\underline{1}75 \times 10^{-2}\ mol$$

$$4.1\underline{1}75 \times 10^{-2}\ \cancel{mol\ O_2} \times \frac{32.00\ g\ O_2}{1\ \cancel{mol\ O_2}} = 1.32\ g\ O_2$$

CHECK The answer is in the correct units. You can quickly check the magnitude of the answer by using molar volume. Under STP 1 L is about 1/22 of 1 mol. Therefore the answer should be about 1/22 the molar mass of oxygen $(1/22 \times 32 = 1.45)$The magnitude of the answer seems reasonable.

FOR PRACTICE 10.11 A common way to make hydrogen gas in the laboratory is to place a metal such as zinc in hydrochloric acid. The hydrochloric acid reacts with the metal to produce hydrogen gas, which is then collected over water. Suppose a student carries out this reaction and collects a total of 154.4 mL of gas at a pressure of 742 mmHg and a temperature of 25 °C. What mass of hydrogen gas (in mg) does the student collect?