$$p_{\text{Argon}} = \frac{n_{\text{Argon}}}{n_{\text{Argon}} + n_{\text{Neon}} + n_{\text{Helium}}} \times 100 = \frac{\dfrac{m_{\text{Argon}}}{M_{\text{Argon}}}}{\dfrac{m_{\text{Argon}}}{M_{\text{Argon}}} + \dfrac{m_{\text{Neon}}}{M_{\text{Neon}}} + \dfrac{m_{\text{Helium}}}{M_{\text{Helium}}}} \times 100$$

$$= \frac{\dfrac{1.20\text{ g}}{39.948\text{ g/mol}}}{\dfrac{1.20\text{ g}}{39.948\text{ g/mol}} + \dfrac{2.60\text{ g}}{20.180\text{ g/mol}} + \dfrac{3.20\text{ g}}{4.0026\text{ g/mol}}} \times 100 = \boxed{3.1\,\%}$$

The percentage of neon is

$$p_{\text{Neon}} = \frac{n_{\text{Neon}}}{n_{\text{Argon}} + n_{\text{Neon}} + n_{\text{Helium}}} \times 100 = \frac{\dfrac{m_{\text{Neon}}}{M_{\text{Neon}}}}{\dfrac{m_{\text{Argon}}}{M_{\text{Argon}}} + \dfrac{m_{\text{Neon}}}{M_{\text{Neon}}} + \dfrac{m_{\text{Helium}}}{M_{\text{Helium}}}} \times 100$$

$$= \frac{\dfrac{2.60\text{ g}}{20.180\text{ g/mol}}}{\dfrac{1.20\text{ g}}{39.948\text{ g/mol}} + \dfrac{2.60\text{ g}}{20.180\text{ g/mol}} + \dfrac{3.20\text{ g}}{4.0026\text{ g/mol}}} \times 100 = \boxed{13.4\,\%}$$

The percentage of helium is

$$p_{\text{Helium}} = \frac{n_{\text{Helium}}}{n_{\text{Argon}} + n_{\text{Neon}} + n_{\text{Helium}}} \times 100 = \frac{\dfrac{m_{\text{Helium}}}{M_{\text{Helium}}}}{\dfrac{m_{\text{Argon}}}{M_{\text{Argon}}} + \dfrac{m_{\text{Neon}}}{M_{\text{Neon}}} + \dfrac{m_{\text{Helium}}}{M_{\text{Helium}}}} \times 100$$

$$= \frac{\dfrac{3.20\text{ g}}{4.0026\text{ g/mol}}}{\dfrac{1.20\text{ g}}{39.948\text{ g/mol}} + \dfrac{2.60\text{ g}}{20.180\text{ g/mol}} + \dfrac{3.20\text{ g}}{4.0026\text{ g/mol}}} \times 100 = \boxed{83.4\,\%}$$

---

9. $\boxed{\text{SSM}}$ ***REASONING*** The number n of moles of water molecules in the glass is equal to the mass m of water divided by the mass per mole. According to Equation 11.1, the mass of water is equal to its density ρ times its volume V. Thus, we have

$$n = \frac{m}{\text{Mass per mole}} = \frac{\rho V}{\text{Mass per mole}}$$

The volume of the cylindrical glass is $V = \pi r^2 h$, where r is the radius of the cylinder and h is its height. The number of moles of water can be written as

$$n = \frac{\rho V}{\text{Mass per mole}} = \frac{\rho\left(\pi r^2 h\right)}{\text{Mass per mole}}$$

SOLUTION The molecular mass of water (H_2O) is 2(1.00794 u) + (15.9994 u) = 18.0 u. The mass per mole of H_2O is 18.0 g/mol. The density of water (see Table 11.1) is $1.00 \times 10^3\text{ kg/m}^3$ or 1.00 g/cm^3.

$$n = \frac{\rho\left(\pi r^2 h\right)}{\text{Mass per mole}} = \frac{\left(1.00\text{ g/cm}^3\right)(\pi)(4.50\text{ cm})^2(12.0\text{ cm})}{18.0\text{ g/mol}} = \boxed{42.4\text{ mol}}$$

---

10. ***REASONING*** The initial solution contains N_0 molecules of arsenic trioxide, and this number is reduced by a factor of 100 with each dilution. After one dilution, in other words, the number of arsenic trioxide molecules in the solution is $N_1 = N_0/100$. After 2 dilutions, the number remaining is $N_2 = N_1/100 = N_0/100^2$. Therefore, after d dilutions, the number N of arsenic trioxide molecules that remain in the solution is given by

$$N = \frac{N_0}{100^d} \qquad (1)$$

We seek the maximum value of d for which at least one molecule of arsenic trioxide remains. Setting $N = 1$ in Equation (1), we obtain $100^d = N_0$. Taking the common logarithm of both sides yields $\log(100^d) = \log N_0$, which we solve for d:

$$d\log 100 = \log N_0 \quad \text{or} \quad 2d = \log N_0 \quad \text{or} \quad d = \tfrac{1}{2}\log N_0 \qquad (2)$$

The initial number N_0 of molecules of arsenic trioxide in the undiluted solution depends upon the number n_0 of moles of the substance present via $N_0 = n_0 N_A$, where N_A is Avogadro's number. Because we know the original mass m of the arsenic trioxide, we can find the number n_0 of moles from the relation $n_0 = \dfrac{m}{\text{Mass per mole}}$. We will use the chemical formula As_2O_3 for arsenic trioxide, and the periodic table found on the inside of the back cover of the textbook to determine the mass per mole of arsenic trioxide.

$$\Delta n = \left[\frac{0.150\ \text{m}^3}{8.31\ \text{J}/(\text{mol}\cdot\text{K})}\right]\left(\frac{1.00\times10^5\ \text{Pa}}{296\ \text{K}} - \frac{9.50\times10^4\ \text{Pa}}{453\ \text{K}}\right) = \boxed{2.31\ \text{mol}}$$

13. ***REASONING*** Since the absolute pressure, volume, and temperature are known, we may use the ideal gas law in the form of Equation 14.1 to find the number of moles of gas. When the volume and temperature are raised, the new pressure can also be determined by using the ideal gas law.

SOLUTION
a. The number of moles of gas is

$$n = \frac{PV}{RT} = \frac{(1.72\times10^5\ \text{Pa})(2.81\ \text{m}^3)}{[8.31\ \text{J}/(\text{mol}\cdot\text{K})][(273.15+15.5)\ \text{K}]} = \boxed{201\ \text{mol}} \qquad (14.1)$$

b. When the volume is raised to 4.16 m^3 and the temperature raised to 28.2 °C, the pressure of the gas is

$$P = \frac{nRT}{V} = \frac{(201\ \text{mol})[8.31\ \text{J}/(\text{mol}\cdot\text{K})][(273.15+28.2)\ \text{K}]}{(4.16\ \text{m}^3)} = \boxed{1.21\times10^5\ \text{Pa}} \qquad (14.1)$$

14. ***REASONING*** According to the ideal gas law, $PV = nRT$, the temperature T is directly proportional to the product PV, for a fixed number n of moles. Therefore, tanks with equal values of PV have the same temperature. Using the data in the table given with the problem statement, we see that the values of PV for each tank are (starting with tank A): $100\ \text{Pa}\cdot\text{m}^3$, $150\ \text{Pa}\cdot\text{m}^3$, $100\ \text{Pa}\cdot\text{m}^3$, and $150\ \text{Pa}\cdot\text{m}^3$. Tanks A and C have the same temperature, while B and D have the same temperature.

SOLUTION The temperature of each gas can be found from the ideal gas law, Equation 14.1:

$$T_A = \frac{P_A V_A}{nR} = \frac{(25.0\ \text{Pa})(4.0\ \text{m}^3)}{(0.10\ \text{mol})[8.31\ \text{J}/(\text{mol}\cdot\text{K})]} = \boxed{120\ \text{K}}$$

$$T_B = \frac{P_B V_B}{nR} = \frac{(30.0\ \text{Pa})(5.0\ \text{m}^3)}{(0.10\ \text{mol})[8.31\ \text{J}/(\text{mol}\cdot\text{K})]} = \boxed{180\ \text{K}}$$

$$T_C = \frac{P_C V_C}{nR} = \frac{(20.0\ \text{Pa})(5.0\ \text{m}^3)}{(0.10\ \text{mol})[8.31\ \text{J}/(\text{mol}\cdot\text{K})]} = \boxed{120\ \text{K}}$$

$$T_D = \frac{P_D V_D}{nR} = \frac{(2.0\ \text{Pa})(75\ \text{m}^3)}{(0.10\ \text{mol})[8.31\ \text{J}/(\text{mol}\cdot\text{K})]} = \boxed{180\ \text{K}}$$

15. ***REASONING AND SOLUTION*** The ideal gas law gives

$$V_2 = \left(\frac{P_1}{P_2}\right)\left(\frac{T_2}{T_1}\right)V_1 = \left(\frac{65.0\ \text{atm}}{1.00\ \text{atm}}\right)\left(\frac{297\ \text{K}}{288\ \text{K}}\right)(1.00\ \text{m}^3) = \boxed{67.0\ \text{m}^3}$$

16. ***REASONING*** We can use the ideal gas law, Equation 14.1 ($PV = nRT$) to find the number of moles of helium in the Goodyear blimp, since the pressure, volume, and temperature are known. Once the number of moles is known, we can find the mass of helium in the blimp.

SOLUTION The number n of moles of helium in the blimp is, according to Equation 14.1,

$$n = \frac{PV}{RT} = \frac{(1.1\times10^5\ \text{Pa})(5400\ \text{m}^3)}{[8.31\ \text{J}/(\text{mol}\cdot\text{K})](280\ \text{K})} = 2.55\times10^5\ \text{mol}$$

According to the periodic table on the inside of the text's back cover, the atomic mass of helium is 4.002 60 u. Therefore, the mass per mole is 4.002 60 g/mol. The mass m of helium in the blimp is, then,

$$m = (2.55\times10^5\ \text{mol})(4.002\ 60\ \text{g/mol})\left(\frac{1\ \text{kg}}{1000\ \text{g}}\right) = \boxed{1.0\times10^3\ \text{kg}}$$

17. ***REASONING*** The maximum number of balloons that can be filled is the volume of helium available at the pressure in the balloons divided by the volume per balloon. The volume of helium available at the pressure in the balloons can be determined using the ideal gas law. Since the temperature remains constant, the ideal gas law indicates that $PV = nRT$ = constant, and we can apply it in the form of Boyle's law, $P_i V_i = P_f V_f$. In this expression V_f is the final volume at the pressure in the balloons, V_i is the volume of the cylinder, P_i is the initial pressure in the cylinder, and P_f is the pressure in the balloons. However, we need to remember that a volume of helium equal to the volume of the cylinder will remain in the cylinder when its pressure is reduced to atmospheric pressure at the point when balloons can no longer be filled.

SOLUTION We have

$$\frac{(N/V)_{\text{Venus}}}{(N/V)_{\text{Earth}}}=\frac{P_{\text{Venus}}/(kT_{\text{Venus}})}{P_{\text{Earth}}/(kT_{\text{Earth}})}=\left(\frac{P_{\text{Venus}}}{P_{\text{Earth}}}\right)\left(\frac{T_{\text{Earth}}}{T_{\text{Venus}}}\right)=\left(\frac{9.0\times10^{6}\text{ Pa}}{1.0\times10^{5}\text{ Pa}}\right)\left(\frac{320\text{ K}}{740\text{ K}}\right)=\boxed{39}$$

Thus, we can conclude that the atmosphere of Venus is 39 times "thicker" than that of Earth.

---

22. ***REASONING*** Pushing down on the pump handle lowers the piston from its initial height $h_i = 0.55$ m above the bottom of the pump cylinder to a final height h_f. The distance d the biker must push the handle down is the difference between these two heights: $d = h_i - h_f$. The initial and final heights of the piston determine the initial and final volumes of air inside the pump cylinder, both of which are the product of the piston height h and its cross-sectional area A:

$$V_i = Ah_i \qquad \text{and} \qquad V_f = Ah_f \tag{1}$$

We will assume that the air in the cylinder may be treated as an ideal gas, and that compressing it produces no change in temperature. Since none of the air escapes from the cylinder up to the point where the pressure inside equals the pressure in the inner tube, we can apply Boyle's law:

$$P_iV_i = P_fV_f \tag{14.3}$$

We will use Equations (1) and Equation 14.3 to determine the distance $d = h_i - h_f$ through which the pump handle moves before air begins to flow. We note that the final pressure P_f is equal to the pressure of air in the inner tube.

SOLUTION Since the cross-sectional area A of the piston does not change, substituting Equations (1) into Equation 14.3 yields

$$P_i\cancel{A}h_i = P_f\cancel{A}h_f \qquad \text{or} \qquad P_ih_i = P_fh_f \tag{2}$$

Solving Equation (2) for the unknown final height h_f, we obtain

$$h_f = \frac{P_ih_i}{P_f} \tag{3}$$

With Equation (3), we can now calculate the distance $d = h_i - h_f$ that the biker pushes down on the handle before air begins to flow from the pump to the inner tube:

$$d = h_i - h_f = h_i - \frac{P_ih_i}{P_f} = h_i\left(1-\frac{P_i}{P_f}\right)$$

$$=(0.55\text{ m})\left(1-\frac{1.0\times10^5\text{ Pa}}{2.4\times10^5\text{ Pa}}\right)=\boxed{0.32\text{ m}}$$

---

23. SSM WWW ***REASONING AND SOLUTION***

a. Since the heat gained by the gas in one tank is equal to the heat lost by the gas in the other tank, $Q_1 = Q_2$, or (letting the subscript 1 correspond to the neon in the left tank, and letting 2 correspond to the neon in the right tank) $cm_1\Delta T_1 = cm_2\Delta T_2$,

$$cm_1(T-T_1) = cm_2(T_2-T)$$

$$m_1(T-T_1) = m_2(T_2-T)$$

Solving for T gives

$$T = \frac{m_2T_2 + m_1T_1}{m_2+m_1} \tag{1}$$

The masses m_1 and m_2 can be found by first finding the number of moles n_1 and n_2. From the ideal gas law, $PV = nRT$, so

$$n_1 = \frac{P_1V_1}{RT_1} = \frac{(5.0\times10^5\text{ Pa})(2.0\text{ m}^3)}{[8.31\text{ J/(mol}\cdot\text{K)}](220\text{ K})} = 5.5\times10^2\text{ mol}$$

This corresponds to a mass $m_1 = (5.5\times10^2\text{ mol})(20.179\text{ g/mol}) = 1.1\times10^4\text{ g} = 1.1\times10^1\text{ kg}$. Similarly, $n_2 = 2.4\times10^2$ mol and $m_2 = 4.9\times10^3$ g = 4.9 kg. Substituting these mass values into Equation (1) yields

$$T = \frac{(4.9\text{ kg})(580\text{ K}) + (1.1\times10^1\text{ kg})(220\text{ K})}{(4.9\text{ kg}) + (1.1\times10^1\text{ kg})} = \boxed{3.3\times10^2\text{ K}}$$

b. From the ideal gas law,

$$P = \frac{nRT}{V} = \frac{[(5.5\times10^2\text{ mol})+(2.4\times10^2\text{ mol})][8.31\text{ J/(mol}\cdot\text{K)}](3.3\times10^2\text{ K})}{2.0\text{ m}^3+5.8\text{ m}^3} = \boxed{2.8\times10^5\text{Pa}}$$

24. ***REASONING*** Since the temperature of the confined air is constant, Boyle's law applies, and $P_{\text{surface}}V_{\text{surface}} = P_h V_h$, where P_{surface} and V_{surface} are the pressure and volume of the air in the tank when the tank is at the surface of the water, and P_h and V_h are the pressure and volume of the trapped air after the tank has been lowered a distance h below the surface of the water. Since the tank is completely filled with air at the surface, V_{surface} is equal to the volume V_{tank} of the tank. Therefore, the fraction of the tank's volume that is filled with air when the tank is a distance h below the water's surface is

$$\frac{V_h}{V_{\text{tank}}} = \frac{V_h}{V_{\text{surface}}} = \frac{P_{\text{surface}}}{P_h}$$

We can find the absolute pressure at a depth h using Equation 11.4. Once the absolute pressure is known at a depth h, we can determine the ratio of the pressure at the surface to the pressure at the depth h.

SOLUTION According to Equation 11.4, the trapped air pressure at a depth $h = 40.0$ m is

$$P_h = P_{\text{surface}} + \rho g h = (1.01\times10^5 \text{ Pa}) + \left[(1.00\times10^3 \text{ kg/m}^3)(9.80 \text{ m/s}^2)(40.0 \text{ m})\right]$$

$$= 4.93\times10^5 \text{ Pa}$$

where we have used a value of $\rho = 1.00\times10^3 \text{ kg/m}^3$ for the density of water. The desired volume fraction is

$$\frac{V_h}{V_{\text{tank}}} = \frac{P_{\text{surface}}}{P_h} = \frac{1.01\times10^5 \text{ Pa}}{4.93\times10^5 \text{ Pa}} = \boxed{0.205}$$

---

25. ***REASONING*** The mass (in grams) of the air in the room is the mass of the nitrogen plus the mass of the oxygen. The mass of the nitrogen is the number of moles of nitrogen times the molecular mass (in grams/mol) of nitrogen. The mass of the oxygen can be obtained in a similar way. The number of moles of each species can be found by using the given percentages and the total number of moles. To obtain the total number of moles, we will apply the ideal gas law. If we substitute the Kelvin temperature T, the pressure P, and the volume V (length × width × height) of the room in the ideal gas law, we can obtain the total number n_{Total} of moles of gas as $n_{\text{Total}} = \frac{PV}{RT}$, because the ideal gas law does not distinguish between types of ideal gases.

SOLUTION Using m to denote the mass (in grams), n to denote the number of moles, and M to denote the molecular mass (in grams/mol), we can write the mass of the air in the room as follows:

$$m = m_{\text{Nitrogen}} + m_{\text{Oxygen}} = n_{\text{Nitrogen}}M_{\text{Nitrogen}} + n_{\text{Oxygen}}M_{\text{Oxygen}}$$

We can now express this result using f to denote the fraction of a species that is present and n_{Total} to denote the total number of moles:

$$m = n_{\text{Nitrogen}}M_{\text{Nitrogen}} + n_{\text{Oxygen}}M_{\text{Oxygen}} = f_{\text{Nitrogen}}n_{\text{Total}}M_{\text{Nitrogen}} + f_{\text{Oxygen}}n_{\text{Total}}M_{\text{Oxygen}}$$

According to the ideal gas law, we have $n_{\text{Total}} = \frac{PV}{RT}$. With this substitution, the mass of the air becomes

$$m = \left(f_{\text{Nitrogen}}M_{\text{Nitrogen}} + f_{\text{Oxygen}}M_{\text{Oxygen}}\right)n_{\text{Total}} = \left(f_{\text{Nitrogen}}M_{\text{Nitrogen}} + f_{\text{Oxygen}}M_{\text{Oxygen}}\right)\frac{PV}{RT}$$

The mass per mole for nitrogen (N_2) and for oxygen (O_2) can be obtained from the periodic table on the inside of the back cover of the text. They are, respectively, 28.0 and 32.0 g/mol. The temperature of 22 °C must be expressed on the Kelvin scale as 295 K (see Equation 12.1). The mass of the air is, then,

$$m = \left(f_{\text{Nitrogen}}M_{\text{Nitrogen}} + f_{\text{Oxygen}}M_{\text{Oxygen}}\right)\frac{PV}{RT}$$

$$= \left[0.79(28.0 \text{ g/mol}) + 0.21(32.0 \text{ g/mol})\right]\frac{(1.01\times10^5 \text{ Pa})\left[(2.5 \text{ m})(4.0 \text{ m})(5.0 \text{ m})\right]}{\left[8.31 \text{ J/(mol}\cdot\text{K)}\right](295 \text{ K})}$$

$$= \boxed{5.9\times10^4 \text{ g}}$$

---

26. ***REASONING*** The ideal gas law is $PV = nRT$. Since the number of moles is constant, this equation can be written as $\frac{PV}{T} = nR = \text{constant}$. Thus, the value of $\frac{PV}{T}$ is the same initially and finally, and we can write

$$\frac{P_0V_0}{T_0} = \frac{P_fV_f}{T_f} \qquad (1)$$

This expression can be solved for the final temperature T_f.

We have no direct data for the initial and final pressures. However, we can deal with this by realizing that pressure is the magnitude of the force applied perpendicularly to a surface divided by the area of the surface. Thus, the magnitudes of the forces that the initial and final pressures apply to the piston (and, therefore, to the spring) are given by Equation 11.3 as

$$\underbrace{F_0 = P_0 A}_{\text{Force applied by initial pressure}} \quad \text{and} \quad \underbrace{F_f = P_f A}_{\text{Force applied by final pressure}} \tag{2}$$

The forces in Equations (2) are applied to the spring, and the force F_x^{Applied} that must be applied to stretch an ideal spring by an amount x with respect to its unstrained length is given by Equation 10.1 as

$$F_x^{\text{Applied}} = kx \tag{3}$$

where k is the spring constant.

Lastly, we note that the final volume is the initial volume plus the amount by which the volume increases as the spring stretches. The increased volume due to the additional stretching is $A(x_f - x_0)$. Therefore, we have

$$V_f = V_0 + A(x_f - x_0) \tag{4}$$

SOLUTION The final temperature can be obtained by rearranging Equation (1) to show that

$$T_f = \left(\frac{P_f V_f}{P_0 V_0}\right) T_0 \tag{5}$$

Into this result we can now substitute expressions for P_0 and P_f. These expressions can be obtained by using Equations (2) in Equation (3) as follows (and recognizing that, for the initial and final forces, $P_0 A = F_x^{\text{Applied}}$ and $P_f A = F_x^{\text{Applied}}$):

$$\underbrace{P_0 A}_{\text{Force applied to spring by initial pressure}} = kx_0 \quad \text{and} \quad \underbrace{P_f A}_{\text{Force applied to spring by final pressure}} = kx_f \tag{6}$$

Thus, we find that

$$P_0 = \frac{kx_0}{A} \quad \text{and} \quad P_f = \frac{kx_f}{A} \tag{7}$$

Finally, we substitute Equations (7) for the initial and final pressure and Equation (4) for the final volume into Equation (5). With these substitutions Equation (5) becomes

$$T_f = \left(\frac{P_f V_f}{P_0 V_0}\right) T_0 = \frac{\left(\frac{kx_f}{A}\right)\left[V_0 + A(x_f - x_0)\right]T_0}{\left(\frac{kx_0}{A}\right)V_0} = \frac{x_f\left[V_0 + A(x_f - x_0)\right]T_0}{x_0 V_0}$$

$$= \frac{(0.1000\text{ m})\left[6.00\times10^{-4}\text{ m}^3 + (2.50\times10^{-3}\text{ m}^2)(0.1000\text{ m} - 0.0800\text{ m})\right](273\text{ K})}{(0.0800\text{ m})(6.00\times10^{-4}\text{ m}^3)}$$

$$= \boxed{3.70\times10^2\text{ K}}$$

27. SSM ***REASONING*** The graph that accompanies Problem 75 in Chapter 12 can be used to determine the equilibrium vapor pressure of water in the air when the temperature is 30.0 °C (303 K). Equation 12.6 can then be used to find the partial pressure of water in the air at this temperature. Using this pressure, the ideal gas law can then be used to find the number of moles of water vapor per cubic meter.

SOLUTION According to the graph that accompanies Problem 75 in Chapter 12, the equilibrium vapor pressure of water vapor at 30.0 °C is approximately 4250 Pa. According to Equation 12.6,

$$\begin{array}{c}\text{Partial}\\ \text{pressure of}\\ \text{water vapor}\end{array} = \left(\begin{array}{c}\text{Percent relative}\\ \text{humidity}\end{array}\right)\left(\begin{array}{c}\text{Equilibrium vapor pressure of}\\ \text{water at the existing temperature}\end{array}\right)\times\frac{1}{100}$$

$$= \frac{(55)(4250\text{ Pa})}{100} = 2.34\times10^3\text{ Pa}$$

The ideal gas law then gives the number of moles of water vapor per cubic meter of air as

$$\frac{n}{V} = \frac{P}{RT} = \frac{(2.34\times10^3\text{ Pa})}{[8.31\text{ J/(mol}\cdot\text{K)}](303\text{ K})} = \boxed{0.93\text{ mol/m}^3} \tag{14.1}$$

28. ***REASONING AND SOLUTION*** If the pressure at the surface is P_1 and the pressure at a depth h is P_2, we have that $P_2 = P_1 + \rho gh$. We also know that $P_1 V_1 = P_2 V_2$. Then,

$$\frac{V_1}{V_2} = \frac{P_2}{P_1} = \frac{P_1 + \rho gh}{P_1} = 1 + \frac{\rho gh}{P_1}$$