Anesthesiologists use a variety of anaesthetic gases in their work, and require training in the precise delivery and monitoring of such gases.

4. It is very important that pressure regulators not be switched among tanks of various gases, particularly if one of them is oxygen. A pressure regulator for oxygen tanks, used on another tank, such as propane or hydrogen (or vice versa), could cause oxygen to mix with a combustible substance at very high pressure, which would almost certainly cause an explosion. Oxygen regulators are colour coded, and must have a warning attached. Propane tanks are reverse threaded so that an oxygen regulator *cannot* be attached to a propane tank (or vice versa), even if you try.
5. The Donald Duck voice made while exhaling helium is due to the speed of sound in helium, which is much higher than in air. Helium atoms move much faster than air molecules at the same temperature, so the resonant sounds within the vocal tract become distorted. The danger in this practice is not from (inert) helium; it is that inhaling helium deeply and repeatedly can cause oxygen deprivation. To make Duck imitations safe, always make the helium inhalation "shallow," and take several deep breaths of air immediately afterward.

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SECTION 9.3 QUESTIONS

Making Connections

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1. Automobile tires are the most common commercial product containing compressed air. Others include scuba diving tanks and portable marine horns. Many commercial products contain other compressed gases. Examples include propane and butane fuel containers, inert (often nitrogen or argon) propellants in aerosol cans, and natural gas burners in barbecues, fireplaces, and furnaces.
2. The safety hazard of compressed gases is primarily the danger of rupture (explosion) of the container, caused by overheating or failure of some part, resulting in injury to people nearby.
3. There may be dangers inherent in the substance itself; it may be toxic, flammable, or corrosive, for instance.

9.4 THE IDEAL GAS LAW

PRACTICE

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Understanding Concepts

1. Volume of a gas in an air (pneumatic) shock absorber can be reduced by decreasing the amount, by lowering the temperature, or by raising the pressure.
2. An ideal gas would have zero attractive force between particles, and zero particle size. Real gases will be closest to this condition for small entities at very low pressures and very high temperatures, because then their particles will be farthest apart and moving fastest, minimizing the effect of intermolecular forces. The lower the intermolecular attraction — as in helium atoms, for example — the closer the behaviour to "ideal".
3. $n_{CH_4} = ?$

$p = 210\text{ kPa}$

$T = 35.0°C = 308.0\text{ K}$

$v = 500\text{ mL} = 0.500\text{ L}$

$R = 8.31\text{ kPa·L/(mol·K)}$

$$pv = nRT$$

$$n_{CH4} = \frac{pv}{RT}$$

$$= \frac{210\ \cancel{kPa} \times 0.500\ \cancel{L}}{\frac{8.31\ \cancel{kPa}\cdot\cancel{L}}{mol\cdot\cancel{K}} \times 308.0\ \cancel{K}}$$

$$n_{CH4} = 0.0410\ mol = 41.0\ mmol$$

The amount of methane gas present in the sample is 41.0 mmol.

4. $m = 50\ kg$

$p = 150\ kPa$

$T = 125°C = 398\ K$

$v_{O_2} = ?$

$M = 32.00\ g/mol$

$R = 8.31\ kPa\cdot L/(mol\cdot K)$

$$n_{O_2} = 50\ k\cancel{g} \times \frac{1\ mol}{32.00\ \cancel{g}}$$

$$n_{O_2} = 1.56\ kmol$$

$$pv = nRT$$

$$v_{O_2} = \frac{nRT}{p}$$

$$= \frac{1.56\ k\cancel{mol} \times \frac{8.31\ \cancel{kPa}\cdot L}{\cancel{mol}\cdot\cancel{K}} \times 398\ \cancel{K}}{150\ \cancel{kPa}}$$

$$v_{O_2} = 34\ kL \text{ or } 34\ m^3$$

The volume of oxygen gas is 34 kL or 34 m^3.

5. (a) $m_{C_3H_8} = ?$

$p = 96.7\ kPa$

$T = 22°C = 295\ K$

$v = 1.00\ L$ (assumed)

$M = 44.11\ g/mol$

$R = 8.31\ kPa\cdot L/(mol\cdot K)$

$$pv = nRT$$

$$n_{C_3H_8} = \frac{pv}{RT}$$

$$= \frac{96.7\ \cancel{kPa} \times 1.00\ \cancel{L}}{\frac{8.31\ \cancel{kPa}\cdot\cancel{L}}{mol\cdot\cancel{K}} \times 295\ \cancel{K}}$$

$$n_{C_3H_8} = 0.0394\ mol$$

$$m_{C_3H_8} = 0.0394\ \cancel{mol} \times \frac{44.11\ g}{1\ \cancel{mol}}$$

$$m_{C_3H_8} = 1.74\ g$$

The mass of 1.00 L is 1.74 g, so the density of propane gas is 1.74 g/L at the conditions specified.

(b) Since propane is denser than air, leaking propane will flow downward and collect in low areas, causing a severe explosion hazard. Propane-powered vehicles are not normally permitted in inside or underground parking lots, for this reason.

6. Since $M = \frac{n}{m}$, and $n = \frac{pv}{RT}$, by substituting we arrive at:

$M = \frac{pv}{mRT}$ for any gas that behaves approximately like an ideal gas.

Applying Inquiry Skills

7. (a) **Analysis**

$M_{gas} = ?$

$m = 1.25$ g

$p = 100$ kPa

$T = 0°C = 273$ K

$v = 1.00$ L

$R = 8.31$ kPa·L/(mol·K)

$$pv = nRT$$

$$n_{gas} = \frac{pv}{RT}$$

$$= \frac{100\ \cancel{\text{kPa}} \times 1.00\ \cancel{\text{L}}}{\frac{8.31\ \cancel{\text{kPa}}\cdot\cancel{\text{L}}}{\text{mol}\cdot\cancel{\text{K}}} \times 273\ \cancel{\text{K}}}$$

$$n_{gas} = 0.0441 \text{ mol}$$

$$M_{gas} = \frac{1.25 \text{ g}}{0.0441 \text{ mol}}$$

$$M_{gas} = 28.4 \text{ g/mol}$$

The molar mass of the gas sample is 28.4 g/mol.

Making Connections

8. From the expression $pv = nRT$, rearranging gives $n = \frac{pv}{RT}$, which may be interpreted as follows: If pressure and volume of gas remain constant — as is the case in a hot-air balloon that is a specific size and always open to the atmosphere — then the amount of gas contained is inversely proportional to the temperature. Warm air, therefore, will have less gas per unit volume, or lower density, than the same air when cool. This is what makes hot-air balloons buoyant.

9. (a) $\text{density}_{CO_2} = ?$

$M = 44.01$ g/mol

$V_{SATP} = 24.8$ mol/L

$$\text{density}_{CO_2} = \frac{44.01 \text{ g}}{1\ \cancel{\text{mol}}} \times \frac{1\ \cancel{\text{mol}}}{24.8 \text{ L}}$$

$$\text{density}_{CO_2} = 1.78 \text{ g/L}$$

The density of carbon dioxide gas is 1.78 g/L at SATP.

(b) Since carbon dioxide is much denser than air, more will be found close to the floor.

(c) $m_{CO} = ?$

$p = 98$ kPa

$T = 20°C = 293$ K

$v = 1.00$ L (assumed)

$M = 28.01$ g/mol

$$R = 8.31\ \text{kPa·L/(mol·K)}$$

$$pv = nRT$$

$$n_{CO} = \frac{pv}{RT}$$

$$= \frac{98\ \cancel{\text{kPa}} \times 1.00\ \cancel{\text{L}}}{\frac{8.31\ \cancel{\text{kPa}}\cdot\cancel{\text{L}}}{\text{mol}\cdot\cancel{\text{K}}} \times 293\ \cancel{\text{K}}}$$

$$n_{CO} = 0.040\ \text{mol}$$

$$m_{CO} = 0.040\ \cancel{\text{mol}} \times \frac{28.01\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$m_{CO} = 1.1\ \text{g}$$

The mass of 1.00 L is 1.1 g; thus, the density of carbon monoxide gas is 1.1 g/L at the stated conditions.

(d) Carbon monoxide is less dense than air, so a detector should be placed close to the ceiling.

(e) The temperature of the gases in question may be very high, which will reduce their densities noticeably.

Reflecting

10. The word "ideal" refers in most cases to a mental construct — a concept that may be defined but is not measurable — reflecting the limit of thought as applied to something. Most scientific formulas, like $v = d/t$ from physics, reflect an ideal situation (any real speed calculated this way is an approximate average value). Concepts like "pure" when applied to actual chemical substances are similarly "ideal."

SECTION 9.4 QUESTIONS

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Understanding Concepts

1. Any real gas has small interparticle attractive forces. These forces will pull the particles together in a condensed state when the particle motion is slow enough (temperature is low enough).
2. Pressure of a gas sample is determined by its amount (directly), its temperature (directly), and its volume (inversely).
3. $n = 30\ \text{mol}$

$$p_{air} = ?$$

$$T = 40°\text{C} = 313\ \text{K}$$

$$v = 50\ \text{L}$$

$$R = 8.31\ \text{kPa·L/(mol·K)}$$

$$pv = nRT$$

$$p_{air} = \frac{nRT}{v}$$

$$= \frac{30\ \cancel{\text{mol}} \times \frac{8.31\ \text{kPa}\cdot\cancel{\text{L}}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}} \times 313\ \cancel{\text{K}}}{50\ \cancel{\text{L}}}$$

$$p_{air} = 1.6 \times 10^3\ \text{kPa or } 1.6\ \text{MPa}$$

The pressure of air in the cylinder is 1.6 MPa.

4. $m = 10.5\ \text{g}$

$$p = 85.0\ \text{kPa}$$

$$T_{NH_3} = ?$$

$$v = 30.0\ \text{L}$$

M = 17.04 g/mol

R = 8.31 kPa·L/(mol·K)

$$n_{NH_3} = 10.5\ \cancel{g} \times \frac{1\ \text{mol}}{17.04\ \cancel{g}}$$

$$n_{NH_3} = 0.616\ \text{mol}$$

$$pv = nRT$$

$$T_{NH_3} = \frac{pv}{nR}$$

$$= \frac{85.0\ \cancel{\text{kPa}} \times 30.0\ \cancel{\text{L}}}{0.616\ \cancel{\text{mol}} \times \dfrac{8.31\ \cancel{\text{kPa}}\cdot\cancel{\text{L}}}{\cancel{\text{mol}}\cdot\text{K}}}$$

$$T_{NH_3} = 498\ \text{K} = 225\ °\text{C}$$

The temperature of the ammonia in the container is 225 °C.

5. n = 1.00 mol

p = 100 kPa (exactly) $= \dfrac{100}{101.325}$ atm (exactly)

T = 298 K (exactly)

v = 24.8 L

R_{gas} = ?

$$pv = nRT$$

$$R_{gas} = \frac{pv}{nT}$$

$$= \frac{\dfrac{100}{101.325}\ \text{atm} \times 24.8\ \text{L}}{1.00\ \text{mol} \times 298\ \text{K}}$$

$$R_{gas} = 0.0821\ \text{atm}\cdot\text{L/(mol}\cdot\text{K)}$$

The value of the gas constant to three significant digits, calculated from the evidence given, is 0.0821 atm·L/(mol·K).

6. (a) m = 1.49 g

p = 117 kPa

T = 42.0°C = 315.0 K

v = 981 mL = 0.981 L

M_{gas} = ?

R = 8.31 kPa·L/(mol·K)

$$pv = nRT$$

$$n_{gas} = \frac{pv}{RT}$$

$$= \frac{117\ \cancel{\text{kPa}} \times 0.981\ \cancel{\text{L}}}{\dfrac{8.31\ \cancel{\text{kPa}}\cdot\cancel{\text{L}}}{\text{mol}\cdot\cancel{\text{K}}} \times 315.0\ \cancel{\text{K}}}$$

$$n_{gas} = 0.0438\ \text{mol}$$

$$M_{gas} = \frac{1.49\ \text{g}}{0.0438\ \text{mol}}$$

$$M_{gas} = 34.0 \text{ g/mol}$$

The molar mass of the gas sample is 34.0 g/mol.

(b) If the compound has the formula XH_3, where H represents hydrogen, then the atomic molar mass of "X" must be (34.0 − 3(1.01)) g/mol, which gives 31.0 g/mol. The atom must be phosphorus, P, which is 31.0 g/mol if rounded to 3 digits, and the gas is therefore $PH_{3(g)}$.

9.5 AIR QUALITY

PRACTICE

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Understanding Concepts

1. The percentage composition by volume of water vapour in air varies widely.
2. The nitrogen in the atmosphere is converted to soluble compounds (useful to plants) by lightning and by nitrogen-fixing bacteria. During decay of the plants (or the animals that ate them) other bacteria convert the nitrogen back to elemental form and release it to the atmosphere.
3. Use of the internal combustion engine is the most significant human activity contributing nitrogen oxides to the atmosphere.
4. Nitrogen dioxide may be decomposed by ultraviolet (UV) light to produce nitrogen oxide and oxygen.

$$NO_{2(g)} \xrightarrow{\text{UV light}} NO_{(g)} + O_{(g)}$$

The atomic oxygen reacts with molecular oxygen to produce ground-level ozone which is toxic to humans.

$$O_{(g)} + O_{2(g)} \rightarrow O_{3(g)}$$

5. (a) Find the total amount of gases in 240 kL of air at SATP:

$$n_{air} = ?$$
$$p = 100 \text{ kPa}$$
$$T = 25°\text{C} = 298 \text{ K}$$
$$v = 240 \text{ kL}$$
$$R = 8.31 \text{ kPa·L/(mol·K)}$$

$$pv = nRT$$
$$n_{air} = \frac{pv}{RT}$$
$$= \frac{100 \cancel{\text{kPa}} \times 240 \text{ k}\cancel{\text{L}}}{\frac{8.31 \cancel{\text{kPa}}\cdot\cancel{\text{L}}}{\text{mol}\cdot\cancel{\text{K}}} \times 298 \cancel{\text{K}}}$$
$$n_{air} = 9.69 \text{ kmol}$$

Then, using percentage composition values from Table 1:

$$n_{N2} = 9.69 \text{ kmol} \times 0.7808$$
$$= 7.57 \text{ kmol}$$
$$n_{O2} = 9.69 \text{ kmol} \times 0.2095$$
$$= 2.03 \text{ kmol}$$
$$n_{Ar} = 9.69 \text{ kmol} \times 0.00934$$

