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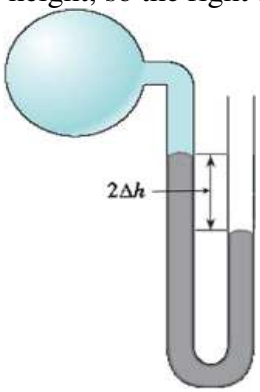
Chapter 2 The Behaviour of Gases

Solutions to Problems in Chapter 2

2.1 A pinhole in the top of the tube would let air leak into the space until the internal pressure matched the external, atmospheric pressure. The barometer would then indicate zero pressure, because the height of the mercury column is determined by the pressure *difference* between inside and outside, and this difference would be zero in the presence of a pinhole.

2.2 The arm of the manometer that feels the greater pressure will display the lower height, so the right-hand side will be lower by 2 Δh.



2.3 When air is pumped into a flat tire, the tire expands and becomes hard. Both these observations are the result of the pressure exerted by the gas molecules making up air.

2.4 The sail of a sailboat feels force from the wind, and this force causes the boat to move across the water. Remember that pressure is defined to be force per unit area, so the force of the wind demonstrates that the gas molecules making up air exert pressure when they strike the sail.

2.5 Useful pressure conversion factors are 1 atm = 1.01325×10^5 Pa, 1 atm = 760 Torr, and 1 bar = 10^5 pascals (Pa).

(a) $455 \text{ Torr}\left(\frac{1 \text{ atm}}{760 \text{ Torr}}\right)\left(\frac{1.01325 \times 10^5 \text{ Pa}}{1 \text{ atm}}\right) = 6.07 \times 10^4 \text{ Pa} = 0.607 \text{ bar}$

(b) $2.45 \text{ atm}\left(\frac{1.01325 \times 10^5 \text{ Pa}}{1 \text{ atm}}\right) = 2.48 \times 10^5 \text{ Pa} = 2.48 \text{ bar}$

(c) $0.46 \text{ Torr}\left(\frac{1 \text{atm}}{760 \text{ atm}}\right)\left(\frac{1.01325 \times 10^5 \text{ Pa}}{1 \text{ atm}}\right) = 61 \text{ Pa} = 6.1 \times 10^{-4} \text{ bar}$

(d) $1.33 \times 10^{-3}\ \text{atm}\left(\dfrac{1.01325 \times 10^{5}\ \text{Pa}}{1\ \text{atm}}\right) = 1.35 \times 10^{2}\ \text{Pa} = 1.35 \times 10^{-3}\ \text{bar}$

2.6 The SI unit for pressure is the pascal: 1 atm = 1.01325×10^5 Pa, and the atmosphere–Torr conversion factor is 1 atm = 760 Torr.

(a) $1.00\ \text{Pa}\left(\dfrac{1\ \text{atm}}{1.01325 \times 10^{5}\ \text{Pa}}\right)\left(\dfrac{760\ \text{Torr}}{1\ \text{atm}}\right) = 7.50 \times 10^{-3}\ \text{Torr}$

(b) $125.6\ \text{bar}\left(\dfrac{10^{5}\ \text{Pa}}{1\ \text{bar}}\right)\left(\dfrac{1\ \text{atm}}{1.01325 \times 10^{5}\ \text{Pa}}\right)\left(\dfrac{760\ \text{Torr}}{1\ \text{atm}}\right) = 9.421 \times 10^{4}\ \text{Torr}$

(c) $75.0\ \text{atm}\left(\dfrac{760\ \text{Torr}}{1\ \text{atm}}\right) = 5.70 \times 10^{4}\ \text{Torr}$

(d) $4.55 \times 10^{-10}\ \text{atm}\left(\dfrac{760\ \text{Torr}}{1\ \text{atm}}\right) = 3.46 \times 10^{-7}\ \text{Torr}$

2.7 The first part of this question asks about number of moles. Rearrange the ideal gas equation to solve for n, and then substitute the appropriate values and do the calculation:

$$n = \frac{pV}{RT} = \frac{(5.00\ \text{bar})(20.0\ \text{L})}{(0.08314\ \text{L bar mol}^{-1}\ \text{K}^{-1})(298\ \text{K})} = 4.04\ \text{mol}$$

The second part of the question asks about the volume if the pressure was different. We can calculate this volume using n and the ideal gas equation, but we can also do the calculation by using proportionalities.

$$p_\text{i}V_\text{i} = p_\text{f}V_\text{f}, \text{from which } V_\text{i} = \frac{p_\text{f}V_\text{f}}{p_\text{i}} = \frac{(5.00\ \text{bar})(20.0\ \text{L})}{1.00\ \text{bar}} = 100.\ \text{L}$$

2.8 The first part of this question asks about number of moles. Rearrange the ideal gas equation to solve for n, and then substitute the appropriate values and do the calculation:

$$n = \frac{pV}{RT} = \frac{(6.47\ \text{bar})(565\ \text{mL})(10^{-3}\ \text{L/mL})}{(0.08314\ \ \text{L bar mol}^{-1}\ \text{K}^{-1})(21.7 + 273.15\ \text{K})} = 0.149\ \text{mol}$$

The second part of the question asks about the volume if the pressure was different. We can calculate this volume using n and the ideal gas equation, but we can also do the calculation by using proportionalities.

$$p_\text{i}V_\text{i} = p_\text{f}V_\text{f}, \text{ from which } V_\text{i} = \frac{(6.47\ \text{bar})(0.565\ \text{L})}{(1.01\ \text{bar})} = 3.62\ \text{L}$$

2.9 When some variables are held fixed, rearrange the ideal gas equation, $pV = nRT$, to

collect fixed values on the right.

(a) *n, R, V* are constant: $\frac{p}{T} = \frac{nR}{V} = \text{constant}; \frac{p_i}{T_i} = \frac{p_f}{T_f}$

(b) $V = \frac{nRT}{p}$

(c) *n, R, T* are constant: $pV = nRT = \text{constant}$; $p_iV_i = p_fV_f$

2.10 To solve for one (or more) variables, rearrange the ideal gas equation, $pV = nRT$, to collect the desired variable(s) on the left.

(a) $n = \frac{pV}{RT}$

(b) *p, R, T* are constant: $\frac{V}{n} = \frac{RT}{p} = \text{constant}$, giving $\frac{V_i}{n_i} = \frac{V_f}{n_f}$

(c) Divide both sides of the equation from (a) by *V*: $\frac{n}{V} = \frac{p}{RT}$

2.11 When some conditions change but others remain fixed, rearrange the ideal gas equation so the constant terms are grouped on the right. In this problem, *n* and *p* are fixed:

$$\frac{V}{T} = \frac{nR}{p} = \text{constant, so } \frac{V_i}{T_i} = \frac{V_f}{T_f}$$

$V_i = 0.255$ L

Convert temperature to kelvins:

$T_i = 25 + 273.15 = 298$ K $\quad T_f = -15 + 273.15 = 258$ K

$$V_f = \frac{V_iT_f}{T_i} = \frac{(0.255\text{ L})(258\text{ K})}{(298\text{ K})} = 0.221\text{ L}$$

2.12 When some conditions change but others remain fixed, rearrange the ideal gas equation so the constant terms are grouped on the right. In this problem, *n* and *V* are fixed:

$\frac{p}{T} = \frac{nR}{V} = \text{constant, so } \frac{p_i}{T_i} = \frac{p_f}{T_f}$

$p_i = 1.075$ bar

Convert temperature to kelvins:

$$T_i = 25 + 273.15 = 298 \text{ K}$$
$$T_f = -15 + 273.15 = 258 \text{ K}$$
$$p_f = \frac{p_i T_f}{T_i} = \frac{(1.075 \text{ bar})(258 \text{ K})}{(298 \text{ K})} = 0.931 \text{ bar}$$

2.13 The equation is valid only for a gas for which n and T are fixed. (a) n and T are fixed, so the equation is valid. (b) n can change, so the equation is not valid. (c) T changes, so the equation is not valid. (d) The equation is not valid for liquids.

2.14 The equation is valid for a gas whose amount is fixed, provided T is in kelvins. Thus, it could be used for (a) and (c) but not for (b) or (d).

2.15 According to Dalton's law of partial pressures, each gas exerts a pressure equal to the total pressure times its mole fraction. Mole fraction can be found from concentration in ppm:

$$X = \left(\frac{\text{ppm}}{10^6}\right)$$

$$p_{NO_2} = (1.011\,\text{bar})\left(\frac{0.78\ \text{molecules NO}_2}{10^6\ \text{molecules of air}}\right) = 7.9\times10^{-7}\ \text{bar}$$

2.16 First convert ppm or % to a mole fraction, and then to a partial pressure. Remember that standard atmospheric pressure is 1 atm:

$$X_{\text{hydrocarbons}} = \left(\frac{150\ \text{ppm}}{10^6\ \text{ppm}}\right) = 1.5\times10^{-4}$$

$$p_{\text{hydrocarbons}} = X_{\text{hydrocarbons}}p_{\text{total}} = (1.5\times10^{-4})(1\text{atm})\left(\frac{1.01325\ \text{bar}}{1\ \text{atm}}\right) = 1.52\times10^{-4}\,\text{bar}$$

$$p_{\text{hydrocarbons}} = 1.52\times10^{-4}\ \text{bar}\left(\frac{10^5\,\text{Pa}}{\text{bar}}\right) = 15.2\ \text{Pa}$$

$$X_{CO} = \left(\frac{0.7\,\text{mol CO}}{100\,\text{mol air}}\right) = 0.007$$

$$p_{CO} = X_{CO}p_{\text{total}} = (0.007)(1\,\text{atm})\left(\frac{1.01325\ \text{bar}}{1\ \text{atm}}\right) = 0.00709\ \text{bar}$$

$$p_{CO} = X_{CO}p_{\text{total}} = (0.00709\ \text{bar})\left(\frac{10^5\,\text{Pa}}{\text{bar}}\right) = 709\ \text{Pa}$$

2.17 According to Dalton's law of partial pressures, each gas exerts a pressure equal to its mole fraction times the total pressure: $p_i = X_i p_{\text{total}}$. Standard atmospheric conditions correspond to $p_{\text{total}} = 1$ atm.

$$p_{N_2} = (0.7808)(1\ \text{atm})\left(\frac{101325\ Pa}{1\ \text{atm}}\right) = 79120\ \text{Pa}$$

$$p_{O_2} = (0.2095)(1\ \text{atm})\left(\frac{101325\ Pa}{1\ \text{atm}}\right) = 21230\ \text{Pa}$$

$$p_{Ar} = (9.34\times10^{-3})(1\ \text{atm})\left(\frac{101325\ Pa}{1\ \text{atm}}\right) = 946\ \text{Pa}$$

$$p_{CO_2} = (3.25 \times 10^{-4})(1 \text{ atm}) \left(\frac{101325\ Pa}{1 \text{ atm}} \right) = 32.9 \text{ Pa}$$

2.18 To calculate partial pressures and mole fractions, first determine the number of moles of each component of the gas mixture and then apply the ideal gas equation.

$$n = \frac{m}{M} \qquad X_i = \frac{n_i}{n_{total}} \qquad p_i = \frac{n_i RT}{V}$$

$$n_{Ar} = 1.25 \text{ g} \left(\frac{1 \text{ mol}}{39.95 \text{ g}} \right) = 3.13 \times 10^{-2} \text{ mol}$$

$$n_{CO} = 1.25 \text{ g} \left(\frac{1 \text{ mol}}{28.01 \text{ g}} \right) = 4.46 \times 10^{-2} \text{ mol}$$

$$n_{CH_4} = 1.25 \text{ g} \left(\frac{1 \text{ mol}}{16.04 \text{ g}} \right) = 7.79 \times 10^{-2} \text{ mol}$$

$$T = 375 + 273.15 = 648 \text{ K}$$

Use the ideal gas equation to determine the partial pressures of each gas:

$$p_{Ar} = \frac{(3.13 \times 10^{-2} \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(648 \text{ K})}{4.00 \text{ L}} = 0.422 \text{ bar}$$

$$p_{CO} = \frac{(4.46 \times 10^{-2} \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(648 \text{ K})}{4.00 \text{ L}} = 0.601 \text{ bar}$$

$$p_{CH_4} = \frac{(7.79 \times 10^{-2} \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(648 \text{ K})}{4.00 \text{ L}} = 1.05 \text{ bar}$$

Obtain the total pressure from the sum of the partial pressures:

p_{total} = 0.422 bar + 0.601 bar + 1.05 bar = 2.07 bar

Divide the moles of each component by the total number of moles:

$n_{total} = (3.13 + 4.46 + 7.79) \times 10^{-2}$ mol = 1.538×10^{-1} mol

$$X_{Ar} = \frac{3.13 \times 10^{-2} \text{ mol}}{1.538 \times 10^{-1} \text{ mol}} = 0.204 \qquad X_{CO} = \frac{4.46 \times 10^{-2} \text{ mol}}{1.538 \times 10^{-1} \text{ mol}} = 0.290$$

$$X_{CH_4} = \frac{7.79 \times 10^{-2} \text{ mol}}{1.538 \times 10^{-1} \text{ mol}} = 0.507$$

2.19 Molecular pictures show the relative numbers of molecules of various substances, which also represent the relative numbers of moles of those substances. The figure contains 8 He atoms and 4 O_2 molecules.

(a) and (b) There are more He atoms, so the pressure due to He and the mole fraction of He are higher.

(c) $X_{He} = \frac{n_{He}}{n_{total}} = \frac{8}{12} = 0.67$

2.20 Molecular pictures show the relative numbers of particles of various substances, which also represent the relative numbers of moles of those substances. Sample I contains 4 A, 3 B, and 3 C, a total of 10 particles; Sample II contains 3 A, 6 B, and 7 C, a total of 16 particles; and Sample III contains 3 A, 5 B, and 18 C, a total of 26 particles.

(a) V and T are the same for all containers, so Sample I, which contains the largest number of particles of A, has the highest value for p_A.

(b) The mole fraction values for B are $X_I = 3/10 = 0.30$, $X_{II} = 6/16 = 0.375$, and $X_{III} = 5/26 = 0.192$, so sample II has the highest value of X_B.

(c) In Sample III, $X_A = 3/26 = 0.115$; ppm = $10^6 X = 1.15 \times 10^5$ ppm.

2.21 To calculate partial pressures from a total pressure, mole fractions are required: $p_i = X_i p_{total}$. Mole fractions can be calculated from the analytical data for the gas sample. The equations needed to solve this problem are

$$n = \frac{m}{M} \qquad X_i = \frac{n_i}{n_{total}} \qquad p_i = \frac{n_i RT}{V}$$

$$n_{CH_4} = 1.57 \text{ g} \left(\frac{1 \text{ mol}}{16.04 \text{ g}} \right) = 9.79 \times 10^{-2} \text{ mol}$$

$$n_{C_2H_6} = 0.41 \text{ g} \left(\frac{1 \text{ mol}}{30.07 \text{ g}} \right) = 1.36 \times 10^{-2} \text{ mol}$$

$$n_{C_3H_8} = 0.020 \text{ g} \left(\frac{1 \text{ mol}}{44.09 \text{ g}} \right) = 4.54 \times 10^{-4} \text{ mol}$$

$n_{total} = 9.79 \times 10^{-2}$ mol + 1.36×10^{-2} mol + 4.54×10^{-4} mol = 1.120×10^{-1} mol. Determine the mole fraction by dividing the moles of each component by the total number of moles:

$$X_{CH_4} = \frac{9.79 \times 10^{-2} \text{ mol}}{1.120 \times 10^{-1} \text{ mol}} = 0.874 \qquad X_{C_2H_6} = \frac{1.36 \times 10^{-2} \text{ mol}}{1.120 \times 10^{-1} \text{ mol}} = 0.121$$

$$X_{C_3H_8} = \frac{4.54 \times 10^{-4} \text{ mol}}{1.120 \times 10^{-1} \text{ mol}} = 4.06 \times 10^{-3}$$

Multiply the mole fractions by the total pressure to obtain the partial pressure (round the final values for ethane and propane to two significant figures, because their masses are only known to two significant figures):

$p_{CH_4} = (0.874)(2.35 \text{ bar}) = 2.05 \text{ bar}$

$p_{C_2H_6} = (0.121)(2.35 \text{ bar}) = 0.28 \text{ bar}$

$p_{C_3H_8} = (4.06 \times 10^{-3})(2.35 \text{ bar}) = 9.5 \times 10^{-3} \text{ bar}$

2.22 To calculate partial pressures from a total pressure, mole fractions are required: $p_i = X_i\, p_{total}$. Mole fractions can be calculated from the analytical data for the gas sample.

$$X = \left(\frac{\text{ppm}}{10^6}\right) = \left(\frac{\text{ppb}}{10^9}\right)$$

$$p_{CO_2} = 113.1 \text{ kPa}\left(\frac{10^3 \text{ Pa}}{1 \text{ kPa}}\right)\left(\frac{487.4 \text{ molecules } CO_2}{10^6 \text{ molecules of air}}\right) = 55.12 \text{ Pa}$$

$$p_{NO} = 113.1 \text{ kPa}\left(\frac{10^3 \text{ Pa}}{1 \text{ kPa}}\right)\left(\frac{10.3 \text{ molecules NO}}{10^9 \text{ molecules of air}}\right) = 1.16 \times 10^{-3} \text{ Pa}$$

$$p_{CO} = 113.1 \text{ kPa}\left(\frac{10^3 \text{ Pa}}{1 \text{ kPa}}\right)\left(\frac{4.2 \text{ molecules CO}}{10^9 \text{ molecules of air}}\right) = 4.8 \times 10^{-4} \text{ Pa}$$

2.23 Stoichiometric calculations require moles and a balanced chemical equation. Balance the equation, determine the number of moles of CO_2 produced, and apply the ideal gas equation to calculate V:

The unbalanced chemical equation is

$C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$

6 C + 12 H + 8 O → 1 C + 2 H + 3 O

Give both CO_2 and H_2O a coefficient of 6 to balance C and H:

$C_6H_{12}O_6 + O_2 \rightarrow 6\ CO_2 + 6\ H_2O$;

6 C + 12 H + 8 O → 6 C + 12 H + 18 O

This leaves 18 O on the product side. 6 O are from glucose, so O_2 needs a coefficient of 6 to balance O and obtain the balanced chemical equation:

$C_6H_{12}O_6\,(g) + 6\ O_2\,(g) \rightarrow 6\ CO_2\,(g) + 6\ H_2O\,(l)$

$M_{glucose} = 6(12.01 \text{ g/mol}) + 12(1.01 \text{ g/mol}) + 6(16.00 \text{ g/mol}) = 180.18 \text{ g/mol}$. Use mass–mole conversions to determine the moles of CO_2 formed:

$$n_{CO_2} = 4.65 \text{ g } C_6H_{12}O_6\left(\frac{1 \text{ mol}}{180.18 \text{ g}}\right)\left(\frac{6 \text{ mol } CO_2}{1 \text{ mol } C_6H_{12}O_6}\right) = 1.548 \times 10^{-1} \text{ mol}$$

$$V = \frac{nRT}{p} = \frac{(1.548 \times 10^{-1}\ \text{mol})(8.314\ \text{L kPa mol}^{-1}\ \text{K}^{-1})(310\ \text{K})}{101.325\ \text{kPa}} = 3.94\ \text{L}$$

2.24 Stoichiometric calculations require moles and a balanced chemical equation. Balance the equation, determine the moles of O_2 produced, and apply the ideal gas equation to calculate V.

The unbalanced equation is

$$KClO_3 \rightarrow KCl + O_2$$

$$1\ K + 1\ Cl + 3\ O \rightarrow 1\ K + 1\ Cl + 2\ O$$

Both K and Cl are already balanced. Balance O by giving O_2 a coefficient of 3/2 and then multiply all coefficients by 2 to eliminate fractions:

$$2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$$

$M_{KClO_3} = 39.10\ \text{g/mol} + 35.45\ \text{g/mol} + 3(16.00\ \text{g/mol}) = 122.55\ \text{g/mol}$ Use mass–mole conversions to determine the moles of CO_2 formed:

$$n_{O_2} = 1.57\ \text{g}\left(\frac{1\ \text{mol}}{122.55\ \text{g}}\right)\left(\frac{3\ \text{mol}\ O_2}{2\ \text{mol}\ KClO_3}\right) = 1.92 \times 10^{-2}\ \text{mol}$$

$$V = \frac{nRT}{p} = \left(\frac{(1.92 \times 10^{-2}\ \text{mol})(0.08314\ \text{L bar mol}^{-1}\ \text{K}^{-1})(298\ \text{K})}{1.020\ \text{bar}}\right) = 0.466\ \text{L}$$

2.25 This is a stoichiometry problem that involves gases. We are asked to determine the final pressure of the container (which is the pressure of chlorine gas). Begin by analyzing the chemistry. The starting materials are Na metal and Cl_2, a gas. The product is NaCl. The balanced chemical reaction is 2 Na (*s*) + Cl_2 (*g*) → 2 NaCl (*s*).

The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the final pressure. Calculations of initial amounts:

$$n_{Na} = 6.90\ \text{g}\left(\frac{1\ \text{mol}}{22.99\ \text{g}}\right) = 0.300\ \text{mol}$$

For chlorine gas, use the ideal gas equation to do pressure–mole conversions. The data must have the same units as those for R:

$$p_{Cl_2} = 1.67 \times 10^5\ Pa\left(\frac{1\ \text{bar}}{10^5\ Pa}\right) = 1.67\ \text{bar}$$

$$n_{Cl_2} = \frac{pV}{RT} = \frac{(1.67\ \text{bar})(3.00\ \text{L})}{(0.08314\ \text{L bar mol}^{-1}\ \text{K}^{-1})(300\ \text{K})} = 0.200\ \text{mol}$$

Divide each initial amount by its coefficient to determine the limiting reactant:

Cl_2: 0.200 mol Na: $\frac{0.300 \text{ mol}}{2} = 0.\ 150 \text{ mol (LR)}$

Use the initial amounts and the balanced equation to construct a table of amounts:

Reaction:	**2 Na (*s*) + Cl₂ (*g*) → 2 NaCl (*s*)**		
Initial amount (mol)	0.300	0.200	0.000
Change (mol)	–0.300	–0.150	+0.300
Final amount (mol)	0.000	0.050	0.300

Use the final amount of Cl_2 and the new temperature (47 °C) to calculate the pressure:

$$p = \frac{nRT}{V} = \frac{(5.0 \times 10^{-2} \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(320 \text{ K})}{3.00 \text{ L}} = 0.443 \text{ bar}$$

2.26 This is a stoichiometry problem that involves gases. We are asked to determine the final partial pressure of each gas. Begin by analyzing the chemistry. The starting materials are CO and O_2, both gases. The product of the reaction is CO_2.
The balanced chemical reaction is: 2 CO (*g*) + O_2 (*g*) → 2 CO_2 (*g*)
The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the partial pressure. Use the ideal gas equation to determine the initial amounts of each gas:

$$n_{CO} = \frac{p_{CO}V}{RT} = \frac{(1.00 \text{ bar})(50.0 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298 \text{ K})} = 2.02 \text{ mol}$$

$$n_{O_2} = \frac{p_{O_2}V}{RT} = \frac{(3.56 - 1.00 \text{ bar})(50.0 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298 \text{ K})} = 5.17 \text{ mol}$$

O_2: 5.17 mol CO: $\frac{2.02 \text{ mol}}{2} = 1.01 \text{ mol}$ (LR)

Use the initial amounts and the balanced equation to construct a table of amounts:

Reaction:	**2 CO (*g*) + O₂ (*g*) → 2 CO₂ (*g*)**		
Initial amount (mol)	2.02	5.17	0.000
Change (mol)	–2.02	–1.01	+2.02
Final amount (mol)	0.00	4.22	2.02

Obtain the partial pressures of each gas using the final amounts from the table:

$p_{CO} = 0$

$$p_{O_2} = \frac{nRT}{V} = \frac{(4.22 \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{(50.0 \text{ L})} = 2.09 \text{ bar}$$

$$p_{CO_2} = \frac{nRT}{V} = \frac{(2.02 \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{(50.0 \text{ L})} = 1.00 \text{ bar}$$

2.27 This is a stoichiometry problem that involves gases. We are asked to determine the mass of product produced. Begin by analyzing the chemistry. The starting materials are N_2 and H_2, both gases. The product of the reaction is NH_3.

The balanced chemical reaction is $N_2 + 3\ H_2 \rightarrow 2\ NH_3$.

The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the partial pressure. Use the ideal gas equation to determine the initial amounts of each gas. The reactor initially contains the gases in 1:1 mole ratio, so

$$p_i = \frac{275 \text{ bar}}{2} = 137.5 \text{ bar for each gas}$$

$$n_i = \frac{p_i V}{RT} = \frac{(137.5 \text{ bar})(8.75 \times 10^3 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(455 + 273.15 \text{ K})} = 1.99 \times 10^4 \text{ mol}$$

Since both gases have the same initial amount, the limiting reactant will be the one with the larger stoichiometric coefficient: H_2 is limiting. Here is the complete amounts table:

Reaction:	**N_2 (*g*)**	**+ 3 H_2 (*g*)**	**→ 2 NH_3 (*g*)**
Initial amount (10^4 mol)	1.99	1.99	0.00
Change (10^4 mol)	$-\frac{1}{3}(1.99)$	–1.99	$+\frac{2}{3}(1.99)$
Final amount (10^4 mol)	1.33	0.00	1.33

Obtain the mass of product formed from the final amount in the table:

$$m_{NH_3} = nM = 1.33 \times 10^4 \text{ mol} \left(\frac{17.04 \text{ g}}{1 \text{ mol}} \right) = 2.27 \times 10^5 \text{ g}$$

2.28 This is a stoichiometry problem that involves gases. We are asked to determine the mass of product produced. Begin by analyzing the chemistry.

The balanced chemical reaction is: $2\ C_2H_4\ (g) + O_2\ (g) \rightarrow 2\ C_2H_4O\ (g)$

The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the partial pressure. Use the ideal gas equation to determine the initial amounts of each gas. The reactor initially contains the gases in 1:1 mole ratio, so

$$p_i = \frac{1.00 \text{ bar}}{2} = 0.500 \text{ bar for each gas}$$

$$n_i = \frac{p_i V}{RT} = \frac{(0.500 \text{ bar})(5.00 \times 10^4 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(280 + 273.15 \text{ K})} = 5.44 \times 10^2 \text{ mol}$$

Since both gases have the same initial amount, the limiting reactant will be the one with the higher stoichiometric coefficient: C_2H_4 is limiting. Here is the complete amounts table:

Reaction:	**2 C_2H_4 (*g*)**	**+**	**O_2 (*g*)**	**→ 2 C_2H_4O (*g*)**
Initial amount (10^2 mol)	5.44		5.44	0.00
Change (10^2 mol)	–5.44		$-\frac{1}{2}$ (5.44)	+5.44
Final amount (10^2 mol)	0		2.72	5.44

Obtain the mass of product formed from the final amount in the table:

$$m_{C_2H_4O} = nM = 5.44 \times 10^2 \text{ mol}\left(\frac{44.05 \text{ g}}{1 \text{ mol}}\right)\left(\frac{10^{-3} \text{ kg}}{1 \text{ g}}\right) = 24.0 \text{ kg}$$

2.29 The yield in a reaction is the ratio of the actual amount produced to the theoretical amount. The calculation in Problem 2.27 gives the theoretical amount:

$$\text{Actual amount} = \text{theoretical amount}\left(\frac{\%\text{ yield}}{100\%}\right)$$

$$\text{Actual amount} = 2.27 \times 10^5 \text{ g}\left(\frac{13\%}{100\%}\right) = 2.9 \times 10^4 \text{ g}$$

2.30 The yield in a reaction is the ratio of the actual amount produced to the theoretical amount. The calculation in Problem 2.28 gives the theoretical amount:

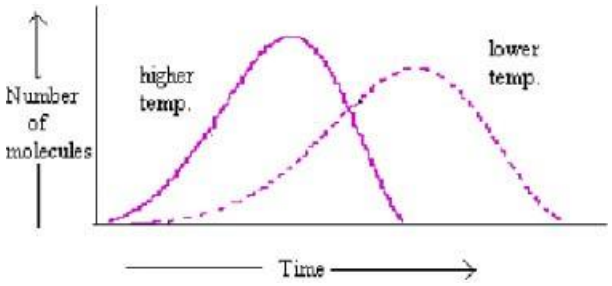
$$\text{Actual amount} = \text{theoretical amount}\left(\frac{\%\text{ yield}}{100\%}\right)$$

$$\text{Actual amount} = 24.0 \text{ kg}\left(\frac{65\%}{100\%}\right) = 16 \text{ kg}$$

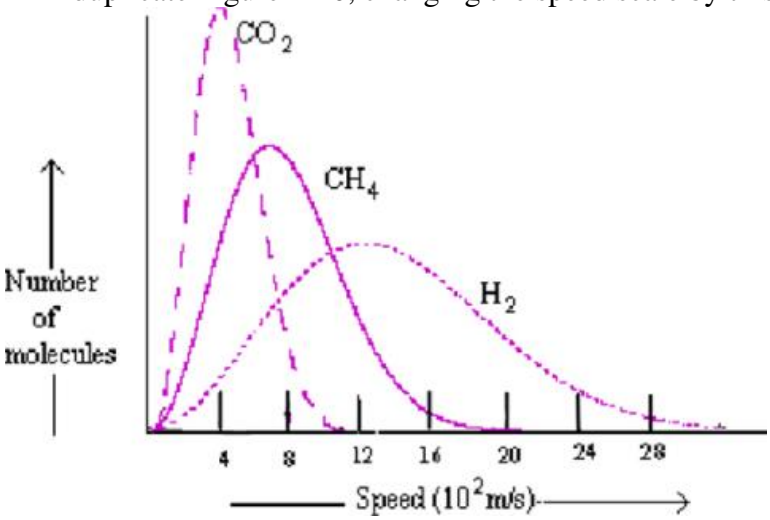
2.31 Molecular speed is related to temperature by the equations, $v = \sqrt{\frac{2E_{\text{kinetic}}}{m}}$, and

$$\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A}, \text{ so } \bar{v} = \sqrt{\frac{3RT}{mN_A}}$$

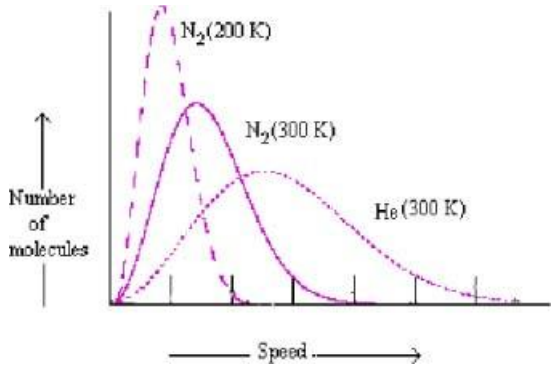
. Thus, molecular speed increases with the square root of the absolute temperature, so molecules will reach the detector sooner at higher temperature, because they have greater speed.



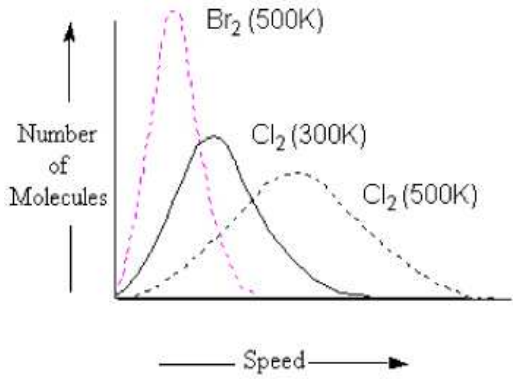
2.32 Molecular speed is related to temperature by the equations, $v = \sqrt{\frac{2E_{\text{kinetic}}}{m}}$, and $\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A} =$, so $\bar{v} = \sqrt{\frac{3RT}{mN_A}}$. Thus, molecular speed varies as the square root of the absolute temperature. At 200 K, the speeds will be smaller than at 300 K by $\sqrt{\frac{200\text{ K}}{300\text{ K}}} = 0.816$. The easiest way to make a new drawing is to duplicate Figure 2-10, changing the speed scale by this factor:



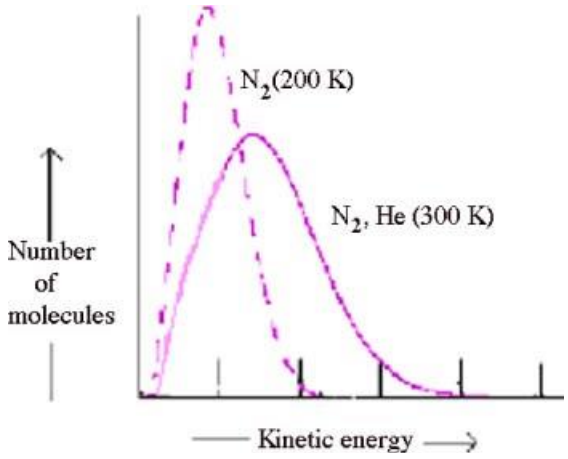
2.33 Molecular speed is related to temperature by two equations: $v = \sqrt{\frac{2E_{\text{kinetic}}}{m}}$, $\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A}$, so $\bar{v} = \sqrt{\frac{3RT}{mN_A}}$. Each distribution of molecular speeds will have the same general shape, with the position of the peak depending on $\sqrt{\frac{T}{m}}$:



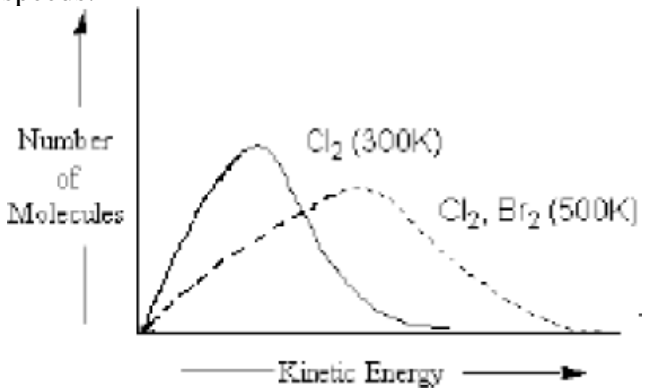
2.34 Molecular speed is related to temperature by the equation, $v = \sqrt{\frac{2E_{\text{kinetic}}}{m}}$, and $\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A} =$, so $\bar{v} = \sqrt{\frac{3RT}{mN_A}}$. Each distribution of molecular speeds will have the same general shape, with the position of the peak depending on $\sqrt{\frac{T}{m}}$:



2.35 Average molecular kinetic energy is directly proportional to temperature and is independent of molecular mass: $\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A}$. N_2 and He at 300 K have identical kinetic energy distributions, whereas the distribution for N_2 at 200 K is shifted to lower energy: The shape of the curve is slightly different for energies than for speeds.



2.36 Average molecular kinetic energy is directly proportional to temperature and is independent of molecular mass: $\bar{E}_{\text{kinetic}} = \dfrac{3RT}{2N_A}$. Br_2 and Cl_2 at 500 K have identical kinetic energy distributions, whereas the distribution for Cl_2 at 300 K is shifted to lower energy: the shape of the curve is slightly different for energies than for speeds.



2.37 The most probable speed at any temperature is related to the most probable kinetic energy by the equation $E_{\text{kinetic}} = \dfrac{1}{2}mv^2$, or $v_{\text{mp}} = \sqrt{\dfrac{2E_{\text{kinetic, mp}}}{m}}$. According to your textbook, the most probable kinetic energy at 300 K is 4.13×10^{-21} J/molecule. Because energy is proportional to temperature, the most probable kinetic energy at any other temperature can be calculated using proportions: $\dfrac{E_{\text{kinetic, 2}}}{E_{\text{kinetic, 1}}} = \dfrac{T_2}{T_1}$. Average kinetic energy, which is not the same as most probable kinetic energy because the distribution of energies is asymmetric, is calculated using $\bar{E}_{\text{kinetic, molar}} = \dfrac{3}{2}RT$ (T in kelvins; 1 J = 1 kg m^2 s^{-2}):

(a) He: $m = \left(\frac{4.00\text{ g}}{1\text{ mol}}\right)\left(\frac{10^{-3}\text{ kg}}{1\text{ g}}\right)\left(\frac{1\text{ mol}}{6.022\times10^{23}\text{ atoms}}\right) = 6.64\times10^{-27}\text{ kg}$

$627\ ^{\circ}\text{C} + 273 = 900\text{ K}$, $E_{\text{kinetic, mp}} = 4.13\times10^{-21}\text{ J}\left(\frac{900\text{ K}}{300\text{ K}}\right) = 1.24\times10^{-20}\text{ J}$

$v_{\text{mp}} = \sqrt{\frac{2(1.24\times10^{-20}\text{ J})}{6.64\times10^{-27}\text{ kg}}} = 1.93\times10^{3}\text{ m/s}$

$\bar{E}_{\text{kinetic, molar}} = \frac{3}{2}\left(8.314\text{ J mol}^{-1}\text{ K}^{-1}\right)(900\text{ K}) = 1.12\times10^{4}\text{ J/mol}$

(b) O_2: $m = \left(\frac{32.00\text{ g}}{1\text{ mol}}\right)\left(\frac{10^{-3}\text{ kg}}{1\text{ g}}\right)\left(\frac{1\text{ mol}}{6.022\times10^{23}\text{ atoms}}\right) = 5.31\times10^{-26}\text{ kg}$

$27^{\circ}\text{C} + 273 = 300\text{ K}$, $E_{\text{kinetic, mp}} = 4.13\times10^{-21}\text{ J}$

$v_{\text{mp}} = \sqrt{\frac{2(4.13\times10^{-21}\text{ J})}{5.31\times10^{-26}\text{ kg}}} = 3.94\times10^{2}\text{ m/s}$

$\bar{E}_{\text{kinetic, molar}} = \frac{3}{2}\left(8.314\text{ J mol}^{-1}\text{ K}^{-1}\right)(300\text{K}) = 3.74\times10^{3}\text{ J/mol}$

(c) SF_6: $M = 32.066\text{ g/mol} + 6(18.998\text{ g/mol}) = 146.1\text{ g/mol}$

$m = \left(\frac{146.1\text{ g}}{1\text{ mol}}\right)\left(\frac{10^{-3}\text{ kg}}{1\text{ g}}\right)\left(\frac{1\text{ mol}}{6.022\times10^{23}\text{ atoms}}\right) = 2.43\times10^{-25}\text{ kg}$

$627^{\circ}\text{C} + 273 = 900\text{ K}$, $E_{\text{kinetic, mp}} = 4.13\times10^{-21}\text{ J}\left(\frac{900\text{K}}{300\text{K}}\right) = 1.24\times10^{-20}\text{J}$

$v_{\text{mp}} = \sqrt{\frac{2(1.24\times10^{-20}\text{ J})}{2.43\times10^{-25}\text{ kg}}} = 3.20\times10^{2}\text{ m/s}$

$\bar{E}_{\text{kinetic, molar}} = \frac{3}{2}\left(8.314\text{ J mol}^{-1}\text{ K}^{-1}\right)(900\text{ K}) = 1.12\times10^{4}\text{ J/mol}$

2.38 The most probable speed at any temperature is related to the most probable kinetic energy by the equation $E_{\text{kinetic}} = \frac{1}{2}mv^2$, or $v_{\text{mp}} = \sqrt{\frac{2E_{\text{kinetic, mp}}}{m}}$. According to your

textbook, the most probable kinetic energy at 300 K is 4.13×10^{-21} J/molecule. Because energy is proportional to temperature, the most probable kinetic energy at any other temperature can be calculated using proportions: $\frac{E_{\text{kinetic, 2}}}{E_{\text{kinetic, 1}}} = \frac{T_2}{T_1}$. Average kinetic energy, which is not the same as most probable kinetic energy because the distribution of energies is asymmetric, is calculated using $\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A}$ per molecule, or $\bar{E}_{\text{kinetic, molar}} = \frac{3}{2}RT$ (T in kelvins; 1 J = 1 kg m^2 s^{-2}):

(a) Ar: $m = \left(\frac{39.948 \text{ g}}{1 \text{ mol}}\right)\left(\frac{10^{-3} \text{ kg}}{1 \text{ g}}\right)\left(\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}}\right) = 6.63 \times 10^{-26} \text{ kg}$

$127 \text{ °C} + 273 = 400 \text{ K}, E_{\text{kinetic, mp}} = 4.13 \times 10^{-21} \text{ J}\left(\frac{400 \text{ K}}{300 \text{ K}}\right) = 5.51 \times 10^{-21} \text{ J}$

$v_{\text{mp}} = \sqrt{\frac{2(5.51 \times 10^{-21} \text{ J})}{6.63 \times 10^{-26} \text{ kg}}} = 4.08 \times 10^2 \text{ m/s}$

$\bar{E}_{\text{kinetic, molar}} = \frac{3}{2}\left(8.314 \text{ J mol}^{-1} \text{ K}^{-1}\right)\left(400 \text{ K}\right) = 4.99 \times 10^3 \text{ J/mol}$

(b) Xe: $m = \left(\frac{131.29 \text{ g}}{1 \text{ mol}}\right)\left(\frac{10^{-3} \text{ kg}}{1 \text{ g}}\right)\left(\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}}\right) = 2.18 \times 10^{-25} \text{ kg}$

$127 \text{ °C} + 273 = 400 \text{ K}, E_{\text{kinetic, mp}} = 4.13 \times 10^{-21} \text{ J}\left(\frac{400 \text{ K}}{300 \text{ K}}\right) = 5.51 \times 10^{-21} \text{ J}$

$v_{\text{mp}} = \sqrt{\frac{2(5.51 \times 10^{-21} \text{ J})}{2.18 \times 10^{-25} \text{ kg}}} = 2.25 \times 10^2 \text{ m/s}$

$\bar{E}_{\text{kinetic, molar}} = \frac{3}{2}\left(8.314 \text{ J mol}^{-1} \text{ K}^{-1}\right)\left(400 \text{ K}\right) = 4.99 \times 10^3 \text{ J/mol}$

(c) C_3H_8: $M = [3(12.01 \text{ g/mol}) + 8(1.008 \text{ g/mol})] = 44.09 \text{ g/mol}$

$m = \left(\frac{44.09 \text{ g}}{1 \text{ mol}}\right)\left(\frac{10^{-3} \text{ kg}}{1 \text{ g}}\right)\left(\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}}\right) = 7.32 \times 10^{-26} \text{ kg}$

$327 \text{ °C} + 273 = 600 \text{ K}, E_{\text{kinetic, mp}} = 4.13 \times 10^{-21} \text{ J}\left(\frac{600 \text{ K}}{300 \text{ K}}\right) = 8.26 \times 10^{-21} \text{ J}$

$v_{\text{mp}} = \sqrt{\frac{2(8.26 \times 10^{-21} \text{ J})}{7.32 \times 10^{-26} \text{ kg}}} = 4.75 \times 10^2 \text{ m/s}$

$\bar{E}_{\text{kinetic, molar}} = \frac{3}{2}\left(8.314 \text{ J mol}^{-1} \text{ K}^{-1}\right)\left(400 \text{ K}\right) = 4.99 \times 10^3 \text{ J/mol}$

2.39 The ideal gas is defined by the conditions that molecular volumes and intermolecular forces both are negligible.

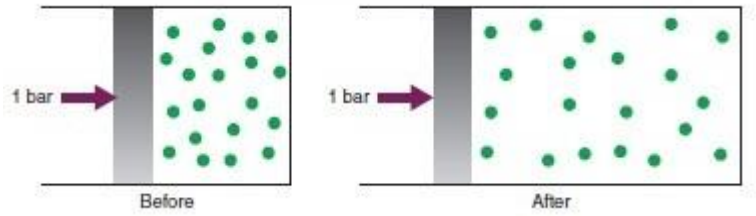
(a) At very high pressure, molecules are very close together, so their volumes are significant compared with the volume of their container; because the first condition is not met, the gas is not ideal. Furthermore, intermolecular forces are significant when the molecules are close together, which influences molecular motion.

(b) At very low temperature, molecules move very slowly, so the forces between molecules, even though small, are sufficient to influence molecular motion; because the second condition is not met, the gas is not ideal.

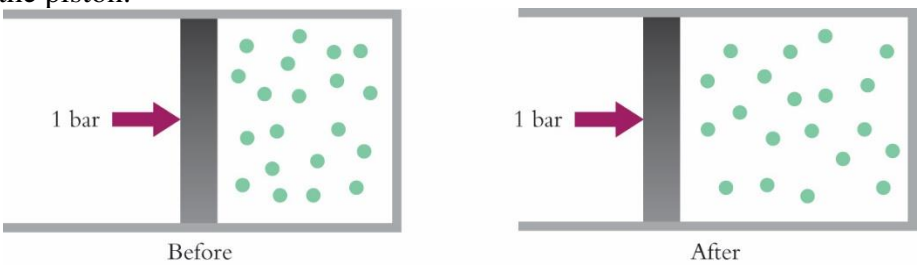
2.40 Pressure is the effect of the impulses due to molecular collisions with the walls.
(a) Pushing the piston in compresses the gas, so the molecular density increases. The frequency of molecular collisions with the walls increases, and so does the pressure.
(b) Removing gas reduces the molecular density. Consequently, the frequency of molecular collisions with the walls decreases, causing the pressure to decrease.
(c) Heating the gas increases the average molecular speed. Consequently, the force exerted per collision and the frequency of collisions with the walls both increase, causing the pressure to increase.

2.41 (a) If the piston is stationary and there is no friction, the forces and pressures on each side must be equal, so the internal pressure is also 1 bar. This pressure is generated by gas molecules colliding with the face of the piston.

(b) When temperature doubles, the increase in average molecular speed results in a doubling of the pressure. The piston will move outward, causing the concentration of gas molecules to decrease. This will lead to a lower frequency of collisions with the wall, reducing the pressure. The piston will stop when the internal pressure is once again 1 bar.



2.42 (a) For the same number of particles and unchanged temperature, the pressure will be $p = \frac{nRT}{V}$. At the molecular level, equal numbers of molecules impart an equal set of impulses to the piston.
(b) When the external pressure is reduced, the piston moves out, reducing the density of gas molecules and thereby reducing the rate of collisions with the face of the piston.



2.43 As a gas cools, its molecules move more slowly, so they impart smaller impulses on the walls of their container. This reduces the internal pressure, so the balloon

collapses until the increase in gas density inside the balloon brings the internal pressure back up to the external pressure of 1.000 bar.

2.44 As a gas cools, its molecules move more slowly, so they impart smaller impulses on the walls of their container. This reduces the pressure exerted by the gas.

2.45 The ideal gas equation, $pV = nRT$, can be used to calculate moles using p–V–T data.

Molar mass is related to moles through $n = \frac{m}{M}$:

$$n = \frac{pV}{RT} = \frac{m}{M} \qquad M = \frac{mRT}{pV}$$

$$T = 25.0 + 273.15 = 298.\ \text{K}$$

Use the modified ideal gas equation to obtain the molar mass:

$$M = \frac{(2.55\ \text{g})(0.08314\ \text{L bar mol}^{-1}\ \text{K}^{-1})(298\ \text{K})}{(0.349\ \text{bar})(1.50\ \text{L})} = 121\ \text{g/mol}$$

The compound contains only C (12.0 g/mol), F (19.0 g/mol), and Cl (35.5 g/mol). The formula can be determined by trial and error. The combination of 1 C, 2 F, and 2 Cl has

$$M = 1(12.0\ \text{g/mol}) + 2(19.0\ \text{g/mol}) + 2(35.5\ \text{g/mol}) = 121\ \text{g/mol}$$

This matches the experimental value. The formula is CF_2Cl_2.

2.46 The ideal gas equation, $pV = nRT$, can be used to calculate moles using p–V–T data.

Also, $n = \frac{m}{M}$. These equations yield two expressions for n:

$$n = \frac{pV}{RT} = \frac{m}{M} \qquad M = \frac{mRT}{pV}$$

Begin by converting the initial data into the units of R:

$$T = 21.5 + 273.15 = 294.7\ \text{K}$$

$$V = 945\ \text{mL}\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right) = 0.945\ \text{L}$$

Use the modified ideal gas equation to obtain the molar mass:

$$M = \frac{(1.63\ \text{g})(0.08314\ \text{L bar mol}^{-1}\ \text{K}^{-1})(294.7\ \text{K})}{(1.50\ \text{bar})(0.945\ \text{L})} = 28.1\ \text{g/mol}$$

Knowing that the compound contains only C and H, find the formula by trial and

error. There are 2 C = 24.02 g/mol, leaving 4 g/mol, which corresponds to 4 H. The molecular formula is C_2H_4.

2.47 Density can be calculated from the ideal gas equation and mole–mass conversions:

$$n = \frac{m}{M} = \frac{pV}{RT} \qquad \frac{m}{V} = \frac{pM}{RT}$$

$$\frac{m}{V} = \frac{pM}{RT} = \frac{(1.007 \text{ bar})(146.05 \text{ g/mol})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(27 + 273 \text{ K})} = 5.90 \text{ g/L}$$

2.48 Density can be calculated from the ideal gas equation and mole–mass conversions:

$$n = \frac{m}{M} = \frac{pV}{RT} \qquad \frac{m}{V} = \frac{pM}{RT}$$

$$\frac{m}{V} = \frac{pM}{RT} = \frac{(0.507 \text{ bar})(2.016 \text{ g/mol})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(-23 + 273 \text{ K})} = 4.91 \times 10^{-2} \text{ g/L}$$

2.49 Molecular speed can be calculated from temperature and molar mass:

$$\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A} \qquad \bar{E}_{\text{kinetic}} = \frac{1}{2} m \bar{v}_{\text{rms}}^2$$

$$\bar{v}^2 = \frac{3RT}{mN_A} = \frac{3RT}{M} \qquad \bar{v} = \left(\frac{3RT}{M} \right)^{1/2}$$

T = 27 °C + 273 = 300. K; M = 32.06 g/mol + 6(18.998 g/mol) = 146.0 g/mol. The mass must be converted to kg/mol for mass units to cancel (1 J = 1 kg m^2 s^{-2}): M = (146.0 g/mol)(10^{-3} kg/g) = 1.460×10^{-1} kg/mol

$$\bar{v}_{\text{rms}} = \left(\frac{3(8.314 \text{ J mol}^{-1} \text{ K}^{-1})(1 \text{ kg m}^2\text{s}^{-2} \text{ /J})(300. \text{ K})}{1.460 \times 10^{-1} \text{ kg/mol}} \right)^{1/2} = 226 \text{ m/s}$$

2.50 Molecular speed can be calculated from temperature and molar mass:

$$\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A} \qquad \bar{E}_{\text{kinetic}} = \frac{1}{2}m\bar{v}_{\text{rms}}^2$$

$$\bar{v}^2 = \frac{3RT}{mN_A} = \frac{3RT}{M} \qquad \bar{v} = \left(\frac{3RT}{M}\right)^{1/2}$$

$T = -23\ ^\circ\text{C} + 273 = 250.\ \text{K}$; $M = 2(1.0078\ \text{g/mol}) = 2.0156\ \text{g/mol}$

The mass must be converted to kg/mol for mass units to cancel ($1\ \text{J} = 1\ \text{kg m}^2\ \text{s}^{-2}$):

$M = (2.0156\ \text{g/mol})(10^{-3}\ \text{kg/g}) = 2.0156 \times 10^{-3}\ \text{kg/mol}$

$$\bar{v} = \left(\frac{3(8.314\ \text{J mol}^{-1}\ \text{K}^{-1})(1\ \text{kg m}^2\text{s}^{-2}/\text{J})(250.\ \text{K})}{2.0156 \times 10^{-3}\ \text{kg/mol}}\right)^{1/2} = 1.76 \times 10^3\ \text{m/s}$$

2.51 Rates of diffusion and effusion depend on the molar masses of the gases because average speeds determine these rates, and these are given by Equation 2-5: $\bar{v}_{\text{rms}} = \left(\frac{3RT}{M}\right)^{1/2}$. For two gases at the same temperature, the one with the smaller molar mass has the faster molecular speed and will diffuse faster. Thus, CH_4 (M = 16 g/mol) diffuses faster through the atmosphere than C_2H_6 (M = 30 g/mol).

2.52 Rates of diffusion and effusion depend on the molar masses of the gases because average speeds determine these rates, and these are given by Equation 2-5: $\bar{v}_{\text{rms}} = \left(\frac{3RT}{M}\right)^{1/2}$. For two gases at the same temperature, the one with the smaller molar mass has the faster molecular speed and will effuse faster. The molar masses of CO and N_2 are the same, however (M = 28 g/mol), so the average molecular speeds of these two gases are the same and they could not be separated by effusion.

2.53 $\dfrac{\text{effusion rate}_A}{\text{effusion rate}_B} = \sqrt{\dfrac{M_B}{M_A}}$ (by Graham's Law)

$$M_{^{14}CO_2} = 14 + 2(16) = 46 \text{ g / mol}$$
$$M_{^{12}CO_2} = 12 + 2(16) = 44 \text{ g / mol}$$
$$\text{thus, } \frac{\text{effusion rate}_{^{14}CO_2}}{\text{effusion rate}_{^{12}CO_2}} = \sqrt{\frac{44 \text{ g / mol}}{46 \text{ g / mol}}} = 0.98$$

$$M_{^{14}CO} = 14 + 16 = 30 \text{ g / mol}$$
$$M_{^{12}CO} = 12 + 16 = 28 \text{ g / mol}$$
$$\text{thus, } \frac{\text{effusion rate}_{^{14}CO}}{\text{effusion rate}_{^{12}CO}} = \sqrt{\frac{28 \text{ g / mol}}{30 \text{ g / mol}}} = 0.97$$

It should be easier to separate ^{14}C from ^{12}C in the form of the monoxide rather than the dioxide, but the values are very close.

2.54 After one diffusion step, the concentration of $^{235}UF_6$ will be 1.0043(0.71%) = 0.713%, after two steps it will be 1.0043(0.713%) = 0.716%, etc. In general, the % of $^{235}UF_6$ after n diffusion steps will therefore be $0.71\%(1.0043)^n$.
After 100 steps, $\%^{235}UF_6 = 0.71\%(1.0043)^{100} = 1.09\%$
After 500 steps, $\%^{235}UF_6 = 0.71\%(1.0043)^{500} = 6.07\%$

2.55 Ideal pressure for F_2 *(g)*:

$$p = \frac{nRT}{V} = \frac{(1.00\,\text{mol})(0.08314\,\text{L bar K}^{-1}\text{mol}^{-1})(298\,\text{K})}{30.0\,\text{L}} = 0.826 \text{ bar}$$

Real pressure for F_2 *(g)*:

$$p = \frac{nRT}{V - nb} - a\left(\frac{n}{V}\right)^2$$
$$= \frac{(1.00 \text{ mol})(0.08314 \text{ L bar K}^{-1}\text{mol}^{-1})(298 \text{ K})}{30.0 \text{ L} - (1.00 \text{ mol})(0.0290 \text{ L / mol})} - (1.171 \text{ bar L}^2 \text{ mol}^{-2})\left(\frac{1.00 \text{ mol}}{30.0 \text{ L}}\right)^2$$
$$= 0.827 \text{ bar} - 0.0013 \text{ bar}$$
$$= 0.826 \text{ bar}$$

The two values are similar because the pressure is low enough that the molecules are far apart and intermolecular forces are small. Also, the temperature is high enough that the kinetic energy of the molecules is high enough to overcome what intermolecular forces do exist.

2.56 Ideal pressure for F_2 *(g)*:

$$p = \frac{nRT}{V} = \frac{(100\ \text{mol})(0.08314\,\text{L}\,bar\,\text{K}^{-1}\,\text{mol}^{-1})(298\,\text{K})}{30.0\ \text{L}} = 82.6\,bar$$

Real pressure for F_2 *(g)*:

$$p = \frac{nRT}{V - nb} - a\left(\frac{n}{V}\right)^2$$

$$= \frac{(100\,\text{mol})(0.08314\,\text{L}\,bar\,\text{K}^{-1}\,\text{mol}^{-1})(298\,\text{K})}{30.0\,\text{L} - (100\,\text{mol})(0.0290\,\text{L/mol})} - (1.171\ bar\,\text{L}^2\,\text{mol}^{-2})\left(\frac{100\,\text{mol}}{30.0\,\text{L}}\right)^2$$

$$= 91.42\ bar - 13.01\ bar$$

$$= 78.4\ bar$$

The real pressure is lower than the ideal pressure because of strong intermolecular forces among F_2 molecules at this high pressure.

2.57 Real pressure of H_2 *(g):*

$$p = \frac{nRT}{V - nb} - a\left(\frac{n}{V}\right)^2$$

$$= \frac{(100\ \text{mol})(0.08314\ \text{L bar}\ \text{K}^{-1}\text{mol}^{-1})(298\ \text{K})}{30.0\ \text{L} - (100\ \text{mol})(0.0266\ \text{L/mol})} - (0.247\ \text{bar L}^2\ \text{mol}^{-2})\left(\frac{100\ \text{mol}}{30.0\,\text{L}}\right)^2$$

$$= 90.62\ \text{bar} - 2.74\ \text{bar}$$

$$= 87.9\ \text{bar}$$

This value is higher than that calculated for fluorine in the previous problem because the intermolecular forces in hydrogen are smaller, because hydrogen is a much smaller molecule than fluorine.

2.58 The van der Waals *a* term for chlorine is much larger than that of fluorine. Likewise, the van der Waals *b* term for chlorine is much larger than that of fluorine. According to the van der Waals equation, a larger *a* term will result in a lower pressure, but a larger *b* term will result in a higher pressure. However, the *a* term has a larger effect on pressure than does the *b* term. Chlorine will therefore have a lower pressure than fluorine under the same conditions.

2.59 Real pressure of O_2 *(g):*

$$p = \frac{nRT}{V - nb} - a\left(\frac{n}{V}\right)^2$$

$$= \frac{(75\text{ mol})(0.08314\text{ L bar K}^{-1}\text{mol}^{-1})(298\text{ K})}{15.0\text{ L} - (75\text{ mol})(0.0318\text{ L/mol})} - (1.378\text{ bar L}^2\text{ mol}^{-2})\left(\frac{75\text{ mol}}{15.0\text{ L}}\right)^2$$

$$= 147.3\text{ bar} - 34.5\text{ bar}$$

$$= 112.8\text{ bar}$$

2.60 The pressure at 700 °C can be calculated using the van der Waals equation:

$$p = \frac{nRT}{V - nb} - a\left(\frac{n}{V}\right)^2$$

$$= \frac{(75\text{ mol})(0.08314\text{ L bar K}^{-1}\text{ mol}^{-1})(700 + 273\text{ K})}{15.0\text{ L} - (75\text{ mol})(0.0318\text{ L/mol})} - (1.378\text{ bar L}^2\text{ mol}^{-2})\left(\frac{75\text{ mol}}{15.0\text{ L}}\right)^2$$

$$= 480.95\text{ bar} - 34.45\text{ bar}$$

$$= 446\text{ bar}$$

This is greater than the maximum pressure for this cylinder, so an explosion is likely.

2.61 Use the ideal gas equation to determine moles, and then do mole–mass–number conversions:

$$n = \frac{pV}{RT} = \left(\frac{(10.0\text{ bar})(725\text{ mL})}{(0.08314\text{ L bar mol}^{-1}\text{ K}^{-1})(925 + 273\text{ K})}\right)\left(\frac{10^{-3}\text{L}}{1\text{ mL}}\right) = 7.28 \times 10^{-2}\text{ mol}$$

$$m = nM = (7.28 \times 10^{-2}\text{ mol})(83.80\text{ g/mol}) = 6.10\text{ g}$$

$$\#\text{ atoms} = nN_A = (7.28 \times 10^{-2}\text{ mol})(6.022 \times 10^{23}\text{ atoms/mol}) = 4.38 \times 10^{22}\text{ atoms}$$

2.62 Use the ideal gas equation to determine moles, and then do mole–mass conversions:

$$n = \frac{pV}{RT} = \frac{(145\text{ bar})(9.50\text{ L})}{(0.08314\text{ L bar mol}^{-1}\text{ K}^{-1})(298\text{ K})} = 55.6\text{ mol}$$

$$m = nM = (55.6\text{ mol})(28.02\text{ g/mol}) = 1.56 \times 10^3\text{ g}$$

2.63 Molecular pictures show the relative numbers of molecules of various substances, which also represent the relative numbers of moles of those substances. Chamber A contains 6 atoms, B contains 12 atoms, and C contains 9 atoms, for a total of 27 atoms, and the chambers have equal volumes and are at the same temperature.

(a) Chamber B contains the most atoms, so it exhibits the highest pressure.

(b) The pressures in the chambers are proportional to the number of atoms:

$$p_A = p_B\left(\frac{\#_A}{\#_B}\right) = 1.0 \text{ bar } \left(\frac{6}{12}\right) = 0.50 \text{ bar}$$

(c) The pressure will increase in proportion to the number of atoms:

$$p_{new} = p_{old}\left(\frac{\#_{new}}{\#_{old}}\right) = 1.0 \text{ bar}\left(\frac{27}{6}\right) = 4.5 \text{ bar}$$

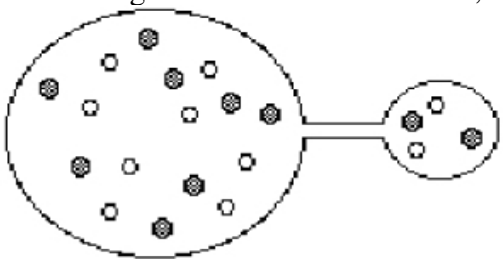
(d) When the valves are opened, the number of atoms in each chamber equalizes:

$$\frac{27}{3} = 9$$

$$p_{new} = p_{old}\left(\frac{\#_{new}}{\#_{old}}\right) = 0.50 \text{ bar}\left(\frac{9}{12}\right) = 0.38 \text{ bar}$$

2.64 Molecular pictures show the relative numbers of molecules of various substances, which also represent the relative numbers of moles of those substances. Tank 1, with a volume of 40 L, contains 10 atoms; tank 2, with a volume of 10 L, also contains 10 atoms.

(a) When the valve is opened, the pressures in the two tanks equalize, and the molecules are distributed in proportion to the volumes. Tank 1 contains four times as many molecules as tank 2, because its volume is four times as large. Thus, tank 1 contains eight molecules of each kind, and tank 2 contains two of each kind:



(b) $p = \dfrac{nRT}{V} = \dfrac{(20 \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(0 + 273 \text{ K})}{(40 \text{ L} + 10 \text{ L})} = 9.1 \text{ bar}$

(c) $p_i = \dfrac{n_i RT}{V} = \dfrac{(10 \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(0 + 273 \text{ K})}{(10 \text{ L})} = 23 \text{ bar}$

The total pressure will therefore be 2(23 bar) = 46 bar

2.65 Much of the information in this problem is not needed, because molecular speed can be calculated from temperature and molar mass:

$$\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A} \qquad \bar{E}_{\text{kinetic}} = \frac{1}{2}m\bar{v}^2$$

$$\bar{v}^2 = \frac{3RT}{mN_A} = \frac{3RT}{M} \qquad \bar{v} = \left(\frac{3RT}{M}\right)^{1/2}$$

The mass must be converted to kg/mol for mass units to cancel (1 J = 1 kg m^2 s^{-2}):

$$M = 3\left(\frac{16.00\text{ g}}{1\text{ mol}}\right)\left(\frac{1\text{ kg}}{1000\text{ g}}\right) = 48.00 \times 10^{-3}\text{ kg/mol}$$

$$\bar{v} = \left(\frac{3(8.314\text{ J mol}^{-1}\text{ K}^{-1})(-25 + 273\text{ K})}{48.00 \times 10^{-3}\text{ kg/mol}}\right)^{1/2} = 359\text{ m/s}$$

2.66 To calculate the pressure, we need the temperature, which can be found from the average kinetic energy: $\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A}$;

$$T = \frac{2N_A E_{\text{kinetic}}}{3R} = \frac{2(6.022 \times 10^{23}\text{ mol}^{-1})(1.02 \times 10^{-20}\text{ J})}{3(8.314\text{ J mol}^{-1}\text{ K}^{-1})} = 493\text{ K}$$

$$n = \frac{m}{M} = \frac{1.55\text{ g}}{39.95\text{ g/mol}} = 3.88 \times 10^{-2}\text{ mol}$$

$$p = \frac{nRT}{V} = \frac{(3.88 \times 10^{-2}\text{ mol})(0.08314\text{ L bar mol}^{-1}\text{ K}^{-1})(493\text{ K})}{5.00\text{ L}} = 0.318\text{ bar}$$

The average speed can be calculated either from the average kinetic energy or from the temperature: $\bar{v} = \left(\frac{3RT}{M}\right)^{1/2}$.

The molar mass must be converted to kg/mol for mass units to cancel: M = 39.95 g/mol or 39.95 × 10^{-3} kg/mol

$$\bar{v} = \left(\frac{3RT}{M}\right)^{1/2} = \sqrt{\frac{3(8.314\text{ J mol}^{-1}\text{ K}^{-1})(493\text{ K})(1\text{ kg m}^2\text{s}^{-2}/\text{J})}{39.95 \times 10^{-3}\text{ kg/mol}}} = 5.55 \times 10^2\text{ m/s}$$

2.67 Use the ideal gas equation to find molar density and Avogadro's number to convert to molecular density:

$$\frac{n}{V} = \frac{p}{RT} = \frac{10^{-10}\text{ bar}}{(0.08314\text{ L bar mol}^{-1}\text{ K}^{-1})(310\text{ K})} = 4 \times 10^{-12}\text{ mol/L}$$

$$\frac{\#}{V} = N_A\left(\frac{n}{V}\right) = \left(\frac{6.022 \times 10^{23}\text{ molecules}}{1\text{ mol}}\right)\left(\frac{4 \times 10^{-12}\text{ mol}}{1\text{ L}}\right) = 2 \times 10^{12}\text{ molecules/L}$$

2.68 Use Avogadro's number to convert to a molar density and the ideal gas equation to calculate p:

$$\frac{n}{V} = \left(\frac{10^{10}\ \text{molecules}}{1\ \text{m}^3}\right)\left(\frac{1\ \text{mol}}{6.022\ \times\ 10^{23}\ \text{molecules}}\right)\left(\frac{10^{-2}\ \text{m}}{1\ \text{cm}}\right)^3\left(\frac{10^3\ \text{cm}^3}{1\ \text{L}}\right)$$

$$= 1.7\ \times\ 10^{-17}\ \text{mol/L}$$

$$p = RT\frac{n}{V} = (0.08314\ \text{L bar mol}^{-1}\ \text{K}^{-1})(25\ \text{K})(1.7\times10^{-17}\ \text{mol/L}) = 4\ \times\ 10^{-17}\ \text{bar}$$

(There is one significant figure because molecular density is known to only one significant figure.)

2.69 The molar mass of gaseous oxygen can be determined using the ideal gas law:

$$M\ =\ \frac{mRT}{pV}$$

Pump out a bulb of known volume, weigh the empty bulb, fill the bulb with oxygen at a measured pressure, and weigh again. For monatomic oxygen, this experiment would give M = 16.00 g/mol, whereas it would give M = 32.00 g/mol for the diatomic gas.

2.70 Molar mass can be determined using the ideal gas equation and the relationship between mass and moles: $M = \frac{mRT}{pV}$. Thus, the experiment must include measurements of the mass, temperature, pressure, and volume of the gas sample whose molar mass is to be calculated. Weigh an evacuated gas bulb. Determine the volume of the bulb by filling it with water and weighing it again. Empty, dry, and pump out the bulb, and then fill it with SO_2 gas to a measured pressure. Weigh again to determine m, from which the molar mass of the gas can be determined. If there is a significant amount of S_2O_4 present, the molar mass will be greater than 64 g/mol, the molar mass of pure SO_2.

2.71 Use Dalton's law of partial pressures to determine partial pressures, and then apply the ideal gas equation to calculate masses.

$$X_{C_3H_6} = \frac{1.00}{5.00} = 0.200 \qquad p_{C_3H_6} = 0.200 \text{ bar}$$

$$X_{O_2} = 1.000 - 0.200 = 0.800 \quad p_{O_2} = 0.800 \text{ bar}$$

$$m_{C_3H_6} = \frac{pVM}{RT} = \frac{(0.200 \text{ bar})(2.00 \text{ L})(42.078 \text{ g/mol})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(23.5 + 273.15 \text{ K})} = 0.682 \text{ g}$$

$$m_{O_2} = \frac{pVM}{RT} = \frac{(0.800 \text{ bar})(2.00 \text{ L})(32.00 \text{ g/mol})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(23.5 + 273.15 \text{ K})} = 2.08 \text{ g}$$

2.72 First calculate moles of each gas, and then apply the ideal gas equation.

$$n_{Ar} = 1.00 \text{ g}\left(\frac{1 \text{ mol}}{39.95 \text{ g}}\right) = 2.50 \times 10^{-2} \text{ mol}$$

$$n_{Ne} = 0.0500 \text{ g}\left(\frac{1 \text{ mol}}{20.18 \text{ g}}\right) = 2.48 \times 10^{-3} \text{ mol}$$

$$X_{Ar} = \frac{2.50 \times 10^{-2} \text{ mol}}{(2.50 \times 10^{-2} \text{ mol}) + (2.48 \times 10^{-3} \text{ mol})} = 0.910$$

$$X_{Ne} = \frac{2.48 \times 10^{-3} \text{ mol}}{(2.50 \times 10^{-2} \text{ mol}) + (2.48 \times 10^{-3} \text{ mol})} = 0.0902$$

$$p_{total} = \frac{nRT}{V} = \frac{(2.75 \times 10^{-2} \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(275 \text{ K})}{5.00 \text{ L}} = 0.126 \text{ bar}$$

$$p_{Ar} = (0.910)(0.126 \text{ bar}) = 0.114 \text{ bar}$$

$$p_{Ne} = (0.0902)(0.126 \text{ bar}) = 0.0114 \text{ bar}$$

2.73 All conditions except the volumes are different for the two samples.

(a) The bulb with the larger number of moles will contain more molecules:

$$n_{H_2} = \frac{pV}{RT} = \frac{(2 \text{ bar})V}{(273 \text{ K})R} = 7 \times 10^{-3}\left(\frac{V}{R}\right) \text{ mol}$$

$$n_{O_2} = \frac{pV}{RT} = \frac{(1 \text{ bar})V}{(298 \text{ K})R} = 3 \times 10^{-3}\left(\frac{V}{R}\right) \text{ mol}$$

The bulb of hydrogen contains more molecules.

(b) $$m_{H_2} = (7 \times 10^{-3})\left(\frac{V}{R}\right) \text{ mol}\left(\frac{2.01 \text{ g}}{1 \text{ mol}}\right) = 1 \times 10^{-2}\left(\frac{V}{R}\right) \text{ g}$$

$$m_{O_2} = (3 \times 10^{-3})\left(\frac{V}{R}\right) \text{ mol}\left(\frac{32.00 \text{ g}}{1 \text{ mol}}\right) = 1 \times 10^{-1}\left(\frac{V}{R}\right) \text{ g}$$

The bulb of oxygen contains more mass.

(c) The average kinetic energy of molecules depends only on the temperature of the gas, so the oxygen molecules, being at higher temperature, have greater average kinetic energy.

(d) The average molecular speed depends on $\sqrt{\frac{T}{m}}$ because $\bar{v} = \sqrt{\frac{3RT}{mN_A}}$. The temperature ratio of H_2:O_2 is $\left(\frac{273}{298}\right) = 0.916$, whereas the mass ratio is $\left(\frac{2}{32}\right) = 0.0625$ Thus, $\sqrt{\frac{T}{m}} = \sqrt{\frac{0.916}{0.0625}} = 3.83$, so the hydrogen molecules have greater average speed by this factor.

2.74 (a) Since we know that both bulbs have the same volume, $\frac{n_{Cl_2}}{n_{N_2}} = \frac{p_{Cl_2}T_{N_2}}{p_{N_2}T_{Cl_2}} = \frac{2\text{ bar}(298\text{ K})}{1\text{ bar}(373\text{ K})} = 1.6.$ Therefore, there are 1.6 moles of Cl_2 per mole of N_2, so there are also more molecules of Cl_2.

(b) To determine mass, we must multiply moles by molar mass. Since there are a greater number of moles of Cl_2 gas and Cl_2 has the higher molar mass, it is easy to see that the Cl_2 bulb has more mass.

(c) Using the equation: $\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A}$, the bulb at the higher temperature will have the higher kinetic energy. Again, this bulb is the Cl_2 bulb at 100 °C compared with the nitrogen bulb at 25 °C.

(d) Using the equation $\bar{v} = \left(\frac{3RT}{M}\right)^{1/2}$ we can determine the average molecular speed for each bulb: $1\text{ kg m}^2\text{s}^{-2}/\text{J}$

$$n_{Ar} = 1.00\text{ g}\left(\frac{1\text{ mol}}{39.95\text{ g}}\right) = 2.50 \times 10^{-2}\text{ mol}$$

$$n_{Ne} = 0.0500\text{ g}\left(\frac{1\text{ mol}}{20.18\text{ g}}\right) = 2.48 \times 10^{-3}\text{ mol}$$

$$X_{Ar} = \frac{2.50 \times 10^{-2}\text{ mol}}{(2.50 \times 10^{-2}\text{ mol}) + (2.48 \times 10^{-3}\text{ mol})} = 0.910$$

$$X_{Ne} = \frac{2.48 \times 10^{-3}\text{ mol}}{(2.50 \times 10^{-2}\text{ mol}) + (2.48 \times 10^{-3}\text{ mol})} = 0.0902$$

$$p_{total} = \frac{nRT}{V} = \frac{(2.75 \times 10^{-2}\text{ mol})(0.08314\text{ L bar mol}^{-1}\text{ K}^{-1})(275\text{ K})}{5.00\text{ L}} = 0.126\text{ bar}$$

$$p_{Ar} = (0.910)(0.126\text{ bar}) = 0.114\text{ bar}$$

$$p_{Ne} = (0.0902)(0.126\text{ bar}) = 0.0114\text{ bar}$$

$$\text{Cl}_2\text{: } \bar{v} = \sqrt{\frac{3(8.314\ \text{J mol}^{-1}\ \text{K}^{-1})(373\ \text{K})(1\ \text{kg m}^2\text{s}^{-2}/\text{J})}{70.90\times10^{-3}\ \text{kg/mol}}} = 362\ \text{m/s}$$

$$\text{N}_2\text{: } \bar{v} = \sqrt{\frac{3(8.314\ \text{J mol}^{-1}\ \text{K}^{-1})(298\ \text{K})(1\ \text{kg m}^2\text{s}^{-2}/\text{J})}{28.02\ \times\ 10^{-3}\ \text{kg/mol}}} = 515\ \text{m/s}$$

Therefore, the N_2 bulb has a higher average molecular speed.

2.75 (a) The most probable kinetic energy at 300 K is 4×10^{-21} J.

(b) $E_{\text{kinetic (most probable)}} = \dfrac{m(v_{\text{most probable}})^2}{2}$ (Remember $1\,\text{J} = 1\,\text{kg m}^2\,\text{s}^{-2}$)

$$v_{\text{most probable}} = \sqrt{\frac{2(4\ \times\ 10^{-21}\ \text{J})(6.022\ \times\ 10^{23}\ \text{molecules/mol})}{32.00\ \times\ 10^{-3}\ \text{kg/mol}}} = 4\ \times\ 10^2\ \text{m/s}$$

2.76 (a) The most probable speed is the speed possessed by the largest number of molecules. Reading from the graph, this is 8×10^2 m/s (the peak position cannot be determined to better than one significant figure).

(b) $E_{\text{kinetic}} = \dfrac{1}{2}mv^2$; because 1 J = 1 kg m^2 s^{-2}, molecular mass must be expressed in kg:

$$m = \frac{M}{N_A} = \left(\frac{17.03\ \text{g/mol}}{6.022\ \times\ 10^{23}\ \text{molecules/mol}}\right)\left(\frac{10^{-3}\ \text{kg}}{1\ \text{g}}\right) = 2.83\ \times\ 10^{-26}\ \text{kg}$$

$$E_{\text{kinetic}} = \frac{1}{2}(2.83\ \times\ 10^{-26}\ \text{kg})(8\ \times\ 10^2\ \text{m/s})^2 = 9\ \times\ 10^{-21}\ \text{kg m}^2\ \text{s}^{-2} = 9\ \times\ 10^{-21}\ \text{J}$$

2.77 Each part of this problem represents a change of one or more conditions, for which a rearranged version of the ideal gas law can be used.

(a) T and n are fixed, so $pV = nRT$ = constant and $V_f = \dfrac{p_iV_i}{p_f}$

$$\text{V}_\text{f} = \frac{(0.700\ \text{bar})(3.00\ \text{L})}{1.007\ \text{bar}} = 2.09\ \text{L}$$

(b) T and n are fixed, so $pV = nRT$ = constant and $p_f = \dfrac{p_iV_i}{V_f}$

$$p_\text{f} = \frac{(0.700\ \text{bar})(3.00\ \text{L})}{2.00\ \text{L}} = 1.05\ \text{bar}$$

(c) V and n are fixed, so $\frac{p}{T} = \text{constant}$ and $p_f = \frac{p_i T_f}{T_i}$

$$p_f = \frac{(0.700 \text{ bar})(50.0 + 273.15 \text{ K})}{273.15 \text{ K}} = 0.828 \text{ bar}$$

2.78 Each part of this problem represents a change of one or more conditions, for which a rearranged version of the ideal gas law can be used.

(a) T and n are fixed, so $pV = nRT = \text{constant}$ and $V_f = \frac{p_i V_i}{p_f}$

$$V_f = \frac{(0.429 \text{ bar})(2.00 \text{ L})}{0.700 \text{ bar}} = 1.23 \text{ L}$$

(b) Here, pressure, temperature, and volume are changing. So the rearranged ideal gas law is $p_f = \frac{p_i V_i T_f}{V_f T_i}$

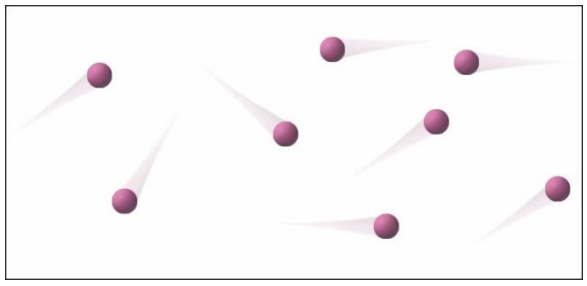
$$P_f = \frac{(0.429 \text{ bar})(2.00 \text{ L})(50.0 + 273.15 \text{ K})}{(1.50 \text{ L})(100.00 + 273.15 \text{ K})} = 0.495 \text{ bar}$$

(c) Here, only pressure and number of moles are changing. So the ideal gas law can be rearranged to $p_f = \frac{p_i n_f}{n_i}$

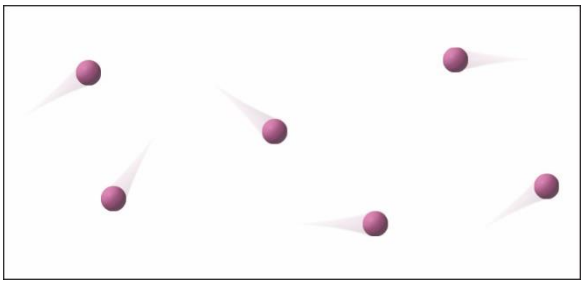
$$p_f = \frac{(0.429 \text{ bar})(1 \text{ mol})}{2 \text{ mol}} = 0.215 \, bar$$

2.79The figure illustrates atomic density and atomic motion of a sample of helium gas. The figure contains eight atoms:

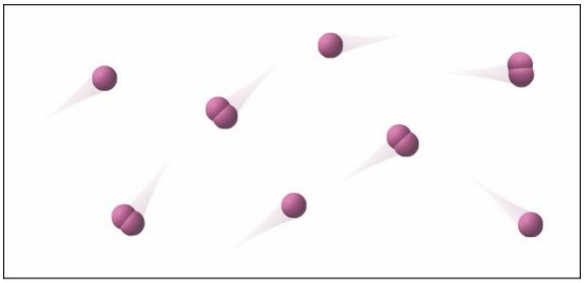
(a) Your drawing should show the same atomic density as the original figure but with longer "tails" on the atoms, indicating that they are moving at higher speeds:



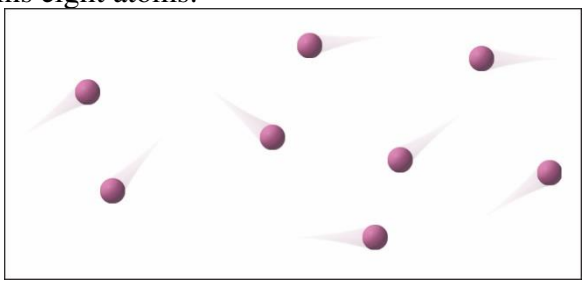
(b) Your drawing should show the same lengths of "tails" as in the original figure, but there should be two fewer atoms:



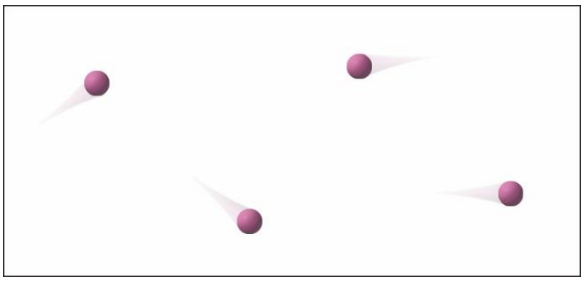
(c) Your drawing should be the same as the original figure except that half the atoms should be replaced with diatomic molecules:



2.80 The figure illustrates atomic density and atomic motion of a sample of helium gas. The figure contains eight atoms:

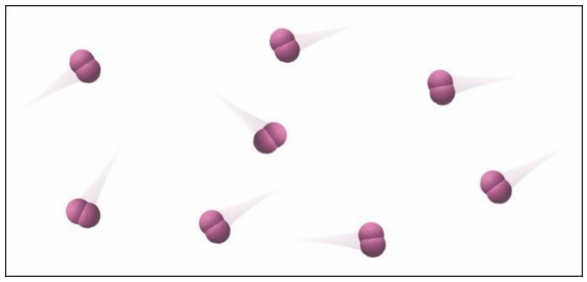


(a) Your drawing should show the same lengths of "tails" as in the original figure (the speed of the molecules is not changing), but there should be half as many atoms.

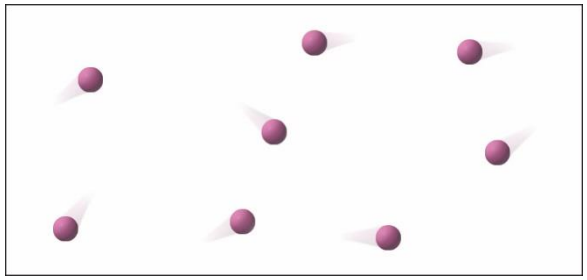


(b) Your drawing should be the same as the original figure except that all the helium

atoms are replaced with nitrogen molecules.



(c) Your drawing should be the same as the original figure except the tails should be shorter on the atoms.



2.81 Use the ideal gas equation, $pV = nRT$, to determine which statements are correct:

(a) At constant T and V, $\frac{p}{n} = \frac{RT}{V} = \text{constant}$. Thus, p is directly proportional to number of moles of gas. The statement is false.

(b) At constant V and n, $\frac{p}{T} = \frac{nR}{V} = \text{constant}$; $\frac{p}{T}$ is constant. The statement is true.

(c) At fixed V and n, $\frac{p}{T} = \frac{nR}{V} = \text{constant}$; $\frac{p}{T}$ is constant; pT is not. The statement is false.

2.82 Use the ideal gas equation, $pV = nRT$, to determine which statements are correct:

(a) False. At fixed n and p, $\frac{V}{T} = \frac{nR}{p} = \text{constant}$. Thus, V is directly proportional to T.

(b) At fixed n and T, $pV = nRT = \text{constant}$. The statement is true.

(c) False. At fixed V and T, $\frac{p}{n} = \frac{RT}{V} = \text{constant}$. Thus, the pressure increases as n increases.

2.83 This is an application of the ideal gas equation to changing gas conditions. The only

constant quantity is n, so $\frac{pV}{T} = nR = \text{constant}$. Therefore, $\frac{p_i V_i}{T_i} = \frac{p_f V_f}{T_f}$.

$p_i = 0.969 \text{ bar} \qquad V_i = 0.963 \text{ L} \qquad p_f = 1.00 \text{ bar}$

$T_i = 22 + 273.15 = 295 \text{ K} \qquad T_f = 15 + 273.15 = 288 \text{ K}$

$$V_f = \left(\frac{p_i V_i}{T_i}\right)\left(\frac{T_f}{p_f}\right) = \frac{(0.969 \text{ bar})(0.963 \text{ L})(288 \text{ K})}{(295 \text{ K})(1.00 \text{ bar})} = 0.911 \text{ L}$$

2.84 There is enough information given to calculate p directly from the ideal gas equation.

$$n = \frac{m}{M} = \frac{96.0 \text{ g}}{32.00 \text{ g/mol}} = 3.00 \text{ mol}$$

$V = 3.00 \text{ L} \qquad T = 27 + 273 = 300. \text{ K}$

$$p = \frac{nRT}{V} = \left(\frac{(3.00 \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(300. \text{ K})}{3.00 \text{ L}}\right) = 29.4 \text{ bar}$$

2.85 Concentrations in parts per billion can be converted to mole fractions, and the partial pressure can then be calculated using $p_i = X_i p_{total}$. To determine molecular concentration, first use the ideal gas equation to calculate n in moles, and then multiply by $\frac{N_A}{V}$:

$X_{SF_6} = (1.0 \text{ ppb})(10^{-9}/\text{ppb}) = 1.0 \times 10^{-9} \qquad p_{SF_6} = (1.0 \times 10^{-9})(1 \text{ bar}) = 1.0 \times 10^{-9} \text{ bar}$

$$V = 1.0 \text{ cm}^3 \left(\frac{10^{-3} \text{ L}}{1 \text{ cm}^3}\right) = 1.0 \times 10^{-3} \text{ L}$$

$$n = \frac{pV}{RT} = \frac{(1.0 \times 10^{-9} \text{ bar})(1.0 \times 10^{-3} \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(21 + 273.15 \text{ K})} = 4.1 \times 10^{-14} \text{ mol}$$

$$\frac{\#}{V} = \frac{(4.1 \times 10^{-14} \text{ mol})(6.022 \times 10^{23} \text{ molecules/mol})}{1 \text{ L}} = 2.5 \times 10^{10} \text{ molecules/L}$$

2.86 Concentrations in parts per million can be converted to mole fractions, and the partial pressure can then be calculated using $p_i = X_i p_{total}$. To determine molecular concentration, first use the ideal gas equation to calculate mol/L, and then multiply by N_A:

$$X_{CO_2} = (385 \text{ ppm})(10^{-6}/\text{ppm}) = 3.85 \times 10^{-4}$$

$$p_{CO_2} = (3.85 \times 10^{-4})(1 \text{ bar}) = 3.85 \times 10^{-4} \text{ bar}$$

$$\frac{n}{V} = \frac{p}{RT} = \frac{3.85 \times 10^{-4} \text{ bar}}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(-45 + 273.15 \text{ K})} = 2.03 \times 10^{-5} \text{ mol/L}$$

$$\frac{\#}{V} = \left(\frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}}\right)\left(\frac{2.03 \times 10^{-5} \text{ mol}}{1.0 \text{ L}}\right) = 1.2 \times 10^{19} \text{ molecules/L}$$

2.87 This is a *P*–*V*–*T* problem. First calculate moles of CO_2, and then use the ideal gas equation to calculate the final pressure:

$$n = \frac{m}{M} = \frac{15.00 \text{ g}}{44.01 \text{ g/mol}} = 0.3408 \text{ mol}$$

$$p_{CO_2} = \frac{(0.3408 \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(273.15 \text{ K})}{0.750 \text{ L}} = 10.3 \text{ bar}$$

2.88 (a) To determine the number of moles of each gas, we only need the information of the individual containers. Moles can be calculated by direct application of the ideal gas equation:

$$n_{N_2} = \frac{pV}{RT} = \frac{(2.0 \text{ bar})(15 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(300. \text{ K})} = 1.2 \text{ mol}$$

$$n_{O_2} = \frac{pV}{RT} = \frac{(3.0 \text{ bar})(1.5 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(300. \text{ K})} = 0.18 \text{ mol}$$

(b) Remember that the pressure of one gas will not affect the pressure of the other. Therefore, treat the expansion as a change of volume at constant *n* and *T*, so *pV* = constant, thus $p_f V_f = p_i V_i$

$$p_{f,\,N_2} = \frac{p_i V_i}{V_f} = \frac{(2.0 \text{ bar})(15 \text{ L})}{(15 + 1.5 \text{ L})} = 1.8 \text{ bar}$$

$$p_{f,\,O_2} = \frac{p_i V_i}{V_f} = \frac{(3.0 \text{ bar})(1.5 \text{ L})}{(15 + 1.5 \text{ L})} = 0.27 \text{ bar}$$

(c) The oxygen molecules will be distributed evenly throughout the volume, so the fraction in the smaller chamber will be the fraction of that chamber's volume relative to the total volume. $\frac{V_{small}}{V_{total}} = \frac{1.5 \text{ L}}{1.5 \text{ L} + 15 \text{ L}} = \frac{1}{11} = 0.091$. There will be 1/11 of the molecules in the smaller container.

2.89 Density can be calculated from the ideal gas equation and mole–mass conversions:

$$n = \frac{m}{M} = \frac{pV}{RT}, \text{ from which } \frac{m}{V} = \frac{pM}{RT}$$

First, calculate the average molar masses of dry air and of moist air:

$$M_{air} = \Sigma_i(X_i M_i)$$

$M_{\text{dry air}} = (0.7808)(28.01 \text{ g/mol}) + (0.2095)(32.00 \text{ g/mol}) + (9.34 \times 10^{-3})(39.948 \text{ g/mol}) = 28.95 \text{ g/mol}$

For moist air at 298 K:

$$p_{H_2O} = 3167 \text{ Pa}\left(\frac{1 \text{ bar}}{10^5 \text{ Pa}}\right) = 3.17 \times 10^{-2} \text{ bar} \qquad X_{H_2O} = 0.0317$$

The total mole fraction of the dry air components is reduced from 1 to (1 – .0317) = 0.9683. Each dry air mole fraction must be multiplied by this factor to obtain the mole fractions in moist air. Thus,

$M_{\text{moist air}} = (0.0317)(18.01 \text{ g/mol}) + (0.9683)(28.95 \text{ g/mol}) = 28.60 \text{ g/mol}$

$$\left(\frac{m}{V}\right)_{\text{dry air}} = \frac{(1 \text{ bar})(28.95 \text{ g/mol})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298 \text{ K})} = 1.17 \text{ g/L}$$

$$\left(\frac{m}{V}\right)_{\text{moist air}} = \frac{(1 \text{ bar})(28.60 \text{ g/mol})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298 \text{ K})} = 1.15 \text{ g/L}$$

2.90 The concentration in parts per million can be converted to a mole fraction, and the partial pressure can then be calculated using $p_i = X_i p_{total}$. To determine molecular concentration, first use the ideal gas equation to calculate mol/L, and then multiply by N_A:

$$X_{O_3} = \frac{0.50 \text{ molecules of } O_3}{10^6 \text{ molecules of air}} = 5.0 \times 10^{-7}$$

$$p_{O_3} = (5.0 \times 10^{-7})(1.016 \text{ bar}) = 5.1 \times 10^{-7} \text{ bar}$$

$$\frac{n}{V} = \frac{p}{RT} = \frac{5.1 \times 10^{-7} \text{ bar}}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(28 + 273)\text{K}} = 2.0 \times 10^{-8} \text{ mol/L}$$

$$\frac{\#}{V} = \left(\frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}}\right)\left(\frac{2.0 \times 10^{-8} \text{ mol}}{1 \text{ L}}\right)\left(\frac{1.00 \text{ L}}{1000 \text{ cm}^3}\right)$$

$$= 1.2 \times 10^{13} \text{ molecules/cm}^3$$

2.91 First, determine moles of liquid oxygen in 150 L. Then, calculate how many moles of dry air contain this amount of oxygen. Finally, use the ideal gas equation to calculate the volume:

$$n = \frac{\rho V}{M} = 150\ \text{L}\left(\frac{10^3\ \text{mL}}{1\ \text{L}}\right)\left(\frac{1.14\ \text{g}}{1\ \text{mL}}\right)\left(\frac{1\ \text{mol}}{32.00\ \text{g}}\right) = 5.34 \times 10^3\ \text{mol}$$

$$n_{\text{air}} = \frac{n_{\text{O}_2}}{X_{\text{O}_2}} = \frac{5.34 \times 10^3\ \text{mol}}{0.2095} = 2.55 \times 10^4\ \text{mol}$$

$$V = \frac{nRT}{p} = \frac{(2.55 \times 10^4\ \text{mol})(0.08314\ \text{L bar mol}^{-1}\ \text{K}^{-1})(298\ \text{K})}{1.000\ \text{bar}} = 6.32 \times 10^5\ \text{L}$$

2.92 This is an empirical formula problem, with *p*–*V*–*T* data added to allow calculation of the molar mass. Use the combustion data to determine the mass percent composition of the compound, and then use elemental molar masses to find the empirical formula:

$$m_{\text{C}} = 363\ \text{mg CO}_2\left(\frac{1\ \text{mol}}{44.01\ \text{g}}\right)\left(\frac{1\ \text{mol C}}{1\ \text{mol CO}_2}\right)\left(\frac{12.01\ \text{g}}{1\ \text{mol}}\right) = 99.1\ \text{mg C}$$

$$\%\ \text{C} = 100\%\left(\frac{99.1\ \text{mg}}{125\ \text{mg}}\right) = 79.3\ \%$$

$$m_{\text{H}} = 63.7\ \text{mg H}_2\text{O}\left(\frac{1\ \text{mol}}{18.02\ \text{g}}\right)\left(\frac{2\ \text{mol H}}{1\ \text{mol H}_2\text{O}}\right)\left(\frac{1.008\ \text{g}}{1\ \text{mol}}\right) = 7.13\ \text{mg H}$$

$$\%\ \text{C} = 100\%\left(\frac{7.13\ \text{mg}}{125\ \text{mg}}\right) = 5.70\%$$

$$\%\ \text{O} = 100\%\ - (79.3\% + 5.70\%) = 15.0\%$$

Assume that the sample weighs 100 g and determine the number of moles of each element in the sample:

$$\text{C: } 79.3\ \text{g}\left(\frac{1\ \text{mol}}{12.01\ \text{g}}\right) = 6.60\ \text{mol C}$$

$$\text{H: } 5.70\ \text{g}\left(\frac{1\ \text{mol}}{1.008\ \text{g}}\right) = 5.65\ \text{mol H}$$

$$\text{O: } 15.0\ \text{g}\left(\frac{1\ \text{mol}}{16.00\ \text{g}}\right) = 0.938\ \text{mol O}$$

Divide each by the smallest among them, 0.938 mol O, to determine the relative amounts of each:

$$\frac{6.60\ \text{mol C}}{0.938\ \text{mol O}} = 7.04\ \text{C/O, round to 7}$$

$$\frac{5.65\ \text{mol H}}{0.938\ \text{mol O}} = 6.02\ \text{H/O, round to 6}$$

The empirical formula is C_7H_6O, with the following molar mass:

$$M = 7(12.01\ \text{g/mol}) + 6(1.01\ \text{g/mol}) + 16.00\ \text{g/mol} = 106\ \text{g/mol}$$

Determine the actual molar mass using the ideal gas law. First convert all properties

of the vapour into the units of *R*:

$$m = 110 \text{ mg}\left(\frac{1 \text{ g}}{10^3 \text{ mg}}\right) = 0.110 \text{ g}$$

$$T = (150 + 273.15 \text{ K}) = 423 \text{ K}$$

$$M = \frac{mRT}{pV} = \frac{(0.110 \text{ g})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(423 \text{ K})}{(0.365 \text{ bar})(0.100 \text{ L})} = 106 \text{ g/mol}$$

The molar mass is the same as the empirical molar mass, so the molecular formula is the same as the empirical formula.

2.93 In addition to a yield problem, information is given about both starting materials, so this is a limiting reactant problem that involves gases. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the theoretical yield. Water can be ignored, because CH_4 is the product of interest. All of the reagents of interest are gases, so pressures can be used as the measures of amounts:

Reaction:	**3 H_2 (*g*)**	**+**	**CO (*g*)**	**→**	**CH_4 (*g*)**
Initial pressure (bar)	20.0		10.0		0.0
Change (bar)	–20.0		–6.67		+6.67
Final pressure (bar)	0.0		3.33		6.67

Use the ideal gas equation and mole–mass conversion to obtain the theoretical yield in grams:

$$n = \frac{pV}{RT} = \frac{(6.67 \text{ bar})(100 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(575 \text{ K})} = 14.0 \text{ mol}$$

Theoretical amount = $nM = (14.0 \text{ mol})(16.04 \text{ g/mol}) = 225 \text{ g}$

Obtain the percent yield by dividing the actual yield by the theoretical yield:

$$\% \text{ Yield} = 100\% \left(\frac{145 \text{ g}}{225 \text{ g}}\right) = 64.4\%$$

2.94 (a) This part is a stoichiometry problem. From the balanced equation, every 2 mole of TNT yields 12 + 5 + 3 = 20 moles of gases. M_{TNT} = 227.14 g/mol:

$$n_{gas} = 1.0 \text{ kg}\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol}}{227 \text{ g}}\right)\left(\frac{20 \text{ mol gas}}{2 \text{ mol TNT}}\right) = 44 \text{ mol}$$

(b) Use the ideal gas equation to obtain the volume:

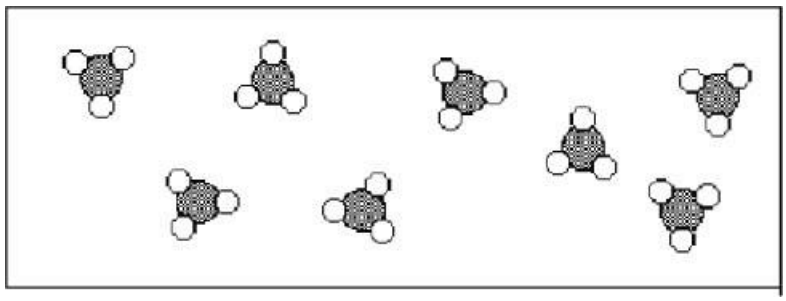
$$V = \frac{(44 \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{(1.00 \text{ bar})} = 1.1 \times 10^3 \text{ L}$$

(c) Determine mole fractions by dividing moles of that gas by the total number of moles. Then use the calculated mole fractions to obtain the partial pressures:

$X_{CO} = (12/20) = 0.60 \qquad X_{H_2} = (5/20) = 0.25 \qquad X_{N_2} = (3/20) = 0.15$

$p_{CO} = (0.60)(1 \text{ bar}) = 0.60 \text{ bar} \qquad p_{H_2} = 0.25 \text{ bar} \qquad p_{N_2} = 0.15 \text{ bar}$

2.95 The chemical reaction is $N_2 + 3\ H_2 \rightarrow 2\ NH_3$. The molecular picture shows 4 moles of N_2 and 12 moles of H_2, which represents stoichiometric proportions. If the reaction goes to completion, all molecules will react to form 8 moles of NH_3:



2.96 This is an example of an ideal gas thermometer. Because n and V are fixed,

$$\frac{p}{T} = \frac{nR}{V} = \text{constant}$$

$$T_f = \frac{T_i p_f}{p_i} = \frac{(273 \text{ K})(0.993 \text{ bar})}{(0.460 \text{ bar})} = 589 \text{ K} \qquad T_f = 589 \text{ K} - 273.15 = 316°\text{C}$$

2.97 The question looks like an empirical formula problem, but the only data provided are p–V–T data for the gaseous substance, which suggests that an ideal gas law calculation can be done. Use the gas data to calculate n, and then use $n = \frac{m}{M}$ to determine M and subsequently determine the formula of the compound:

$$n = \frac{pV}{RT} = \frac{(0.736 \text{ bar})(0.100 \text{ L})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(30 + 273 \text{ K})} = 2.92 \times 10^{-3} \text{ mol}$$

$$M = \frac{m}{n} = \frac{0.500 \text{ g}}{2.92 \times 10^{-3} \text{ mol}} = 171 \text{ g/mol}$$

The compound contains one nickel atom (M = 58.69 g/mol) and x CO molecules (M = 28.01 g/mol), so 171 g/mol = 58.69 g/mol + (x)(28.01 g/mol)

$$x = \frac{171 \text{ g/mol} - 58.69 \text{ g/mol}}{28.01 \text{ g/mol}} = 4 \qquad \text{The formula is } Ni(CO)_4$$

2.98 The only things that are changing in the problem are the number of moles (which is halved), temperature, and pressure. Volume is constant:

$$p_f = \frac{n_f T_f p_i}{n_i T_i} = \frac{(1\text{ mol})(-5+273\text{ K})(0.921\text{ bar})}{(2\text{ mol})(23+273\text{ K})} = 0.417\text{ bar}$$

2.99 At the top of Mount Everest, the atmospheric pressure is 0.333 bar, much lower than the pressure at sea level, 1.01 bar. A lower pressure corresponds to a smaller molecular density. Thus, when people say "the air is thin," they are referring to the fact that air is less dense at higher elevations.

2.100 (a) Each container contains the same number of gas molecules. Because Container A has half the volume as Container B, it will have twice the pressure.

(b) Container A has the higher partial pressure of molecular hydrogen. Container A has five molecules of H_2 and Container B has eight molecules of H_2 but twice the volume. Thus, Container A has a higher number of molecules per unit volume, producing a higher partial pressure of H_2.

(c) The graph shows one large peak and one small peak, indicating an unequal amount of two different species. The lighter gas, molecular hydrogen, will take less time to reach the detector. The first peak is the larger peak, indicating that the container with a greater proportion of molecular hydrogen, Container B, is the one used in the experiment.

2.101 There are two interconnected parts to this problem. Calculate moles of oxygen using the ideal gas equation, and then use stoichiometric analysis for $KClO_3$:

(a) Begin by converting all values into units matching those of R:

$$V = 22.96\text{ mL}\left(\frac{10^{-3}\text{ L}}{1\text{ mL}}\right) = 0.02296\text{ L}$$

$$T = (25.17+273.15)\text{ K} = 298.32\text{ K}$$

$$n = \frac{pV}{RT} = \frac{(1.012\text{ bar})(0.02296\text{ L})}{(0.08314\text{ L bar mol}^{-1}\text{ K}^{-1})(298.32\text{ K})} = 9.370\times10^{-4}\text{ mol}$$

(b) The balanced reaction is: $2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$

$$n_{KClO_3} = (9.370\times10^{-4}\text{ mol O}_2)\left(\frac{2\text{ mol KClO}_3}{3\text{ mol O}_2}\right) = 6.245\times10^{-4}\text{ mol}$$

(c) $m_{KClO_3} = nM = (6.245\times10^{-4}\text{ mol})(122.55\text{ g/mol}) = 7.654\times10^{-2}\text{ g}$

$$\%\ KClO_3 = 100\%\left(\frac{m_{KClO_3}}{m_{total}}\right) = 100\%\left(\frac{7.654\times10^{-2}\text{ g}}{0.1054\text{ g}}\right) = 72.62\%$$

2.102 This is an empirical formula problem, with *p*–*V*–*T* data added to allow calculation of the molar mass. First, determine the empirical formula from mass percentages. Consider 100 g of the compound, which contains the following amounts:

$$\text{C: } 71.22 \text{ g}\left(\frac{1 \text{ mol}}{12.011 \text{ g}}\right) = 5.930 \text{ mol C} \qquad \text{H: } 14.94 \text{ g}\left(\frac{1 \text{ mol}}{1.0079 \text{ g}}\right) = 14.82 \text{ mol H}$$

$$\text{N: } 13.84 \text{ g}\left(\frac{1 \text{ mol}}{14.007 \text{ g}}\right) = 0.9881 \text{ mol N}$$

Divide each by the smallest among them, 0.9881 mol N:

$$\text{C: } \left(\frac{5.930 \text{ mol C}}{0.9881 \text{ mol N}}\right) = 6.001 \text{ mol C/mol N, round to 6 C/N}$$

$$\text{H: } \left(\frac{14.82 \text{ mol H}}{0.9881 \text{ mol N}}\right) = 15.00 \text{ mol H/mol N}$$

Empirical formula: $C_6H_{15}N$, empirical M = 101 g/mol

Next, use the *p*–*V*–*T* data to determine the actual molar mass:

$$m = 250 \text{ mg}\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right) = 0.250 \text{ g}$$

$$V = 150 \text{ mL}\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) = 0.150 \text{ L} \qquad T = 150 + 273 \text{ K} = 423 \text{ K}$$

$$M = \frac{mRT}{pV} = \frac{(0.250 \text{ g})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(423 \text{ K})}{(0.580 \text{ bar})(0.150 \text{ L})} = 101 \text{ g/mol}$$

The molar mass is the same as the empirical molar mass, so the molecular formula is the same as the empirical formula.

2.103 There are two interconnected parts to this problem. Use stoichiometric analysis to calculate the amount of H_2O produced, and then use the ideal gas equation to determine the final pressures:

The reaction is $CuSO_4 \cdot 5H_2O \rightarrow CuSO_4 + 5\ H_2O$

$$M_{CuSO_4 \cdot 5H_2O} = 63.55 \text{ g/mol} + 32.07 \text{ g/mol} + 9(16.00 \text{ g/mol}) + 10(1.01 \text{ g/mol}) = 249.72 \text{ g/mol}$$

$$n_{H_2O} = 2.50 \text{ g } CuSO_4 \cdot 5H_2O\left(\frac{1 \text{ mol}}{249.72 \text{ g}}\right)\left(\frac{5 \text{ mol } H_2O}{1 \text{ mol } CuSO_4 \cdot 5H_2O}\right) = 5.01 \times 10^{-2} \text{ mol}$$

$$p_{H_2O} = \frac{(5.01 \times 10^{-2} \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(227 + 273.15 \text{ K})}{4.00 \text{ L}} = 0.521 \text{ bar}$$

For the air in the vessel, the only change from initial conditions is a temperature change, so

$$p_{\text{f, air}} = \frac{p_{\text{i}} T_{\text{f}}}{T_{\text{i}}} = \frac{(1.00 \text{ bar})(227 + 273.15 \text{ K})}{300 \text{ K}} = 1.67 \text{ bar}$$

$$p_{\text{total}} = 1.67 \text{ bar} + 0.521 \text{ bar} = 2.19 \text{ bar}$$

2.104 The balloon will rise when its net density is less than that of the air that surrounds it. Thus,

$$m_{\text{air}} = m_{\text{He}} + 350 \text{ kg}$$
$$m_{\text{air}} - m_{\text{He}} = 350 \text{ kg}$$

However, the volume of air displaced is the same as the volume of He inside the balloon, so we have:

$$\rho_{\text{air}} V - \rho_{\text{He}} V = 350 \text{ kg}$$
$$V = \frac{350 \text{ kg}}{\rho_{\text{air}} - \rho_{\text{He}}}$$

The molar mass of dry air is calculated from the molecular weights of nitrogen and oxygen, the main constituents of dry air:

$$M_{\text{air}} = \left(\frac{78\%}{100\%}(28.02 \text{ g/mol})\right) + \left(\frac{22\%}{100\%}(32.00 \text{ g/mol})\right) = 28.9 \text{ g/mol}$$

The density of dry air is therefore:

$$\rho_{\text{air}} = \frac{pM}{RT} = \frac{(1 \text{ bar})(28.9 \text{ g/mol})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298)} = 1.17 \text{ g/L}$$

The density of helium gas under these conditions is

$$\rho_{\text{He}} = \frac{pM}{RT} = \frac{(1 \text{ bar})(4.003 \text{ g/mol})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298)} = 0.162 \text{ g/L}$$

And so the volume is

$$V = \frac{350 \text{ kg}}{\rho_{\text{air}} - \rho_{\text{He}}} = \frac{350 \times 10^3 \text{ g}}{(1.17 - 0.162) \text{g/L}} = 3.47 \times 10^5 \text{ L}$$

2.105 A balloon will rise if its net density is less than that of air. For this to be true, the total mass of the balloon plus its payload must be smaller than the mass of the air that it displaces. This gives a mass relationship that must be satisfied:

$m_{\text{cool air}} = m_{\text{hot air}} + 350 \text{ kg}$ $m_{\text{cool air}} - m_{\text{hot air}} = 3.50 \times 10^5 \text{ g}$

Use the ideal gas equation to link m with T, the temperature of the hot air:

$$n = \frac{m}{M} = \frac{pV}{RT} \qquad m = \frac{pVM}{RT}$$

Because all these quantities except for temperature are the same for cool air and hot air, common terms can be collected:

$$\frac{pVM_{\text{air}}}{R}\left(\frac{1}{T_{\text{cool}}} - \frac{1}{T_{\text{hot}}}\right) = 3.50 \times 10^5 \text{ g}$$

The molar mass of air can be found from the mole fractions of N_2 and O_2 (minor components contribute negligibly):

M_{air} = (0.7808)(28.01 g/mol) + (0.2095)(32.00 g/mol) = 28.6 g/mol

The volume of the balloon is the volume found in Problem 2.115: 3.47×10^5 L

$$\frac{(1.00 \text{ bar})(3.47 \times 10^5 \text{ L})(28.6 \text{ g/mol})}{(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})}\left(\frac{1}{T_{\text{cool}}} - \frac{1}{T_{\text{hot}}}\right) = 3.50 \times 10^5 \text{ g}$$

$$\left(\frac{1}{T_{\text{cool}}} - \frac{1}{T_{\text{hot}}}\right) = 0.002932 \text{ K}^{-1} \qquad T_{\text{cool}} = 298 \text{ K}$$

$$\left(\frac{1}{298 \text{ K}} - \frac{1}{T_{\text{hot}}}\right) = 0.002932 \text{ K}^{-1} \qquad T_{\text{hot}} = 2.36 \times 10^3 \text{ K}$$

The material of the balloon would probably not withstand this hot a temperature. Hot air balloons must have much larger volumes than helium balloons per unit payload.

2.106 As the mouse breathes, it inhales air containing oxygen. Some of this oxygen is used for metabolism, resulting in CO_2, which the mouse exhales. The amount of oxygen consumed is therefore equal to the amount of CO_2 exhaled. The solid KOH in the chamber absorbs all the CO_2 and any water the mouse exhales. Any reduction in pressure is therefore due to the removal of oxygen from the atmosphere.

The initial amount of gas in the chamber is

$$n_{\text{initial}} = \frac{p_{\text{initial}}V}{RT} = \frac{1.02 \text{ bar}(2.05 \text{ L})}{(0.08314 \text{ L bar K}^{-1}\text{mol}^{-1})(298 \text{ K})} = 0.08439 \text{ mol}$$

And after two hours, the amount of gas remaining is

$$n_{\text{initial}} = \frac{p_{\text{initial}}V}{RT} = \frac{0.967 \text{ bar}(2.05 \text{ L})}{(0.08314 \text{ L bar K}^{-1}\text{mol}^{-1})(298 \text{ K})} = 0.0800 \text{ mol}$$

The amount of oxygen consumed is therefore 0.08439 – 0.0800 mol = 0.00439 mol.

(0.00439 mol)(32.00 g/mol) = 0.140 g O_2.

2.107 Think about the final conditions for the bottle. It contains air and helium, with $p_{total} = p_{air} + p_{He}$. Each partial pressure must be calculated separately and then summed for the total pressure. The air pressure can be calculated as a variation in conditions. Temperature changes, and although the volume of the *bottle* is fixed, the *gas* volume changes because initially there is 0.100 L of liquid helium present.

Thus, $\frac{pV}{T} = nR = \text{constant}$.

$p_i = 1.0 \text{ bar}$ $V_i = 2.00 \text{ L} - 0.100 \text{ L} = 1.90 \text{ L}$ $T_i = 95 \text{ K}$

$p_f = ?$ $V_f = 2.00 \text{ L}$ $T_f = 25 + 273.15 = 298 \text{ K}$

$$p_{f,\,air} = \left(\frac{p_i V_i}{T_i}\right)\left(\frac{T_f}{V_f}\right) = \frac{(1.0 \text{ bar})(1.90 \text{ L})(298 \text{ K})}{(95 \text{ K})(2.00 \text{ L})} = 2.98 \text{ bar}$$

The pressure due to helium can be calculated by finding the number of moles of liquid helium, and then applying the ideal gas law for the final conditions:

$$n = \frac{m}{M} = \frac{Vd}{M} = \frac{(0.100 \text{ L})(10^3 \text{ mL/L})(0.147 \text{ g/mL})}{(4.00 \text{ g/mol})} = 3.68 \text{ mol}$$

$$p_{He} = \frac{nRT}{V} = \frac{(3.68 \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{(2.00 \text{ L})} = 45.6 \text{ bar}$$

$$p_{total} = 2.98 \text{ bar} + 45.6 \text{ bar} = 48.6 \text{ bar}$$

2.108 Dry air, according to Table 2-4, has a mole fraction of oxygen of 0.2095. The amount of oxygen in the pressurized tank is calculated as follows:

$$n_{O_2} = X_{O_2} n_{air} = X_{O_2}\left(\frac{p_{air} V_{air}}{RT}\right)$$

$$= 0.2095\left(\frac{(165 \text{ bar})(12.5 \text{ L})}{(0.08314 \text{ L bar K}^{-1}\text{mol}^{-1})(27 + 273) \text{ K}}\right) = 82.69 \text{ mol } O_2$$

(82.69 mol)(32.00 g/mol) = 2646 g O_2.

At a rate of 14.0 g O_2 per minute, the tank will last $\frac{2646 \text{ g } O_2}{14.0 \text{ g/min}} = 189 \text{ min}$

Allowing 6 minutes for surfacing, and neglecting the time required to dive to depth, you could therefore stay at depth for 183 minutes.

2.109 This is primarily a stoichiometry problem, but a gas law calculation is needed to determine the volume of gas produced. The reactions are

$S(s) + O_2(g) \rightarrow SO_2(g)$

$SO_2(g) + CaO(s) \rightarrow CaSO_3(s)$

The reaction stoichiometry is in 1:1 mole ratios, so $n_S = n_{SO_2} = n_{CaO} = n_{CaSO_3}$.

Determine the amount (in moles) of sulphur in the coal, which equals all other

amounts:

$$m_S = 1.00 \text{ tonne}\left(\frac{10^3 \text{ kg}}{1 \text{ tonne}}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{4.55\%}{100\%}\right) = 4.55 \times 10^4 \text{ g S}$$

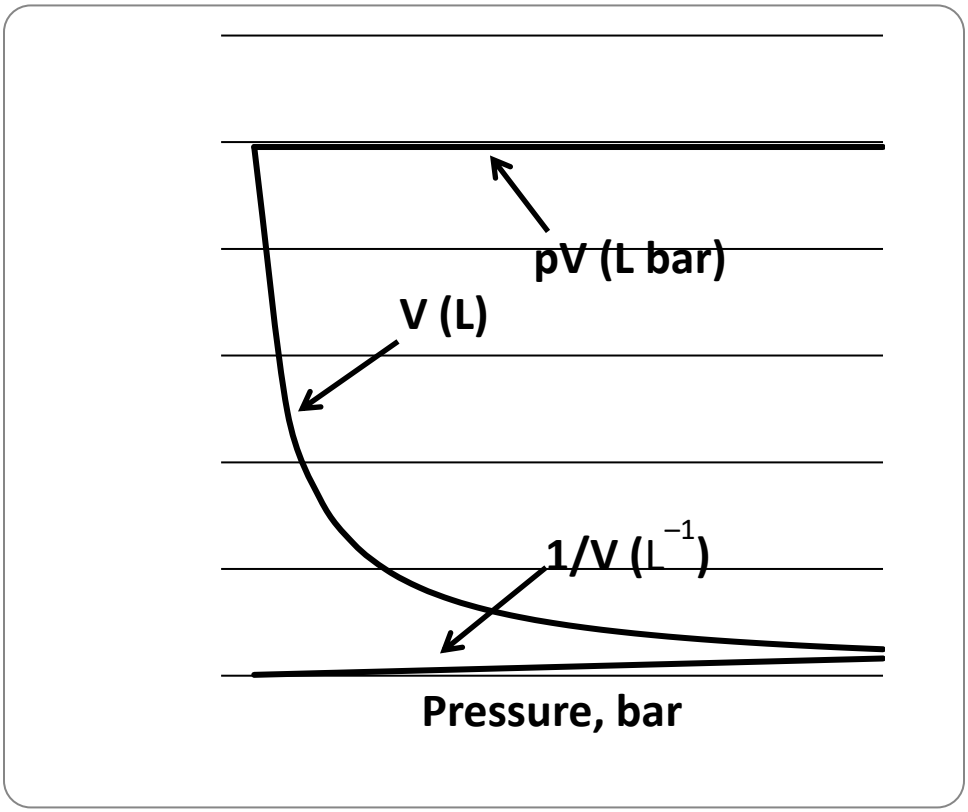
$$n = \frac{m}{M} = \frac{4.55 \times 10^4 \text{ g}}{32.07 \text{ g/mol}} = 1.42 \times 10^3 \text{ mol S}$$

$$V = \frac{(1.42 \times 10^3 \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(55 + 273.15 \text{ K})}{1.00 \text{ bar}} = 3.87 \times 10^4 \text{ L}$$

$$m_{CaO} = nM = (1.42 \times 10^3 \text{ mol})(56.08 \text{ g/mol}) = 7.96 \times 10^4 \text{ g}$$

$$m_{CaSO_3} = nM = (1.42 \times 10^3 \text{ mol})(120.2 \text{ g/mol}) = 1.71 \times 10^5 \text{ g}$$

2.110



The plot shows that as p increases, V decreases. This is explained by the ideal gas equation, $V = \frac{nRT}{p}$.

The plot of pV versus p is linear and horizontal. According to the ideal gas equation, $pV = nRT$. As long as n and T are constant, the product pV will also be constant. This is also referred to as Boyle's Law.

The plot of $1/V$ versus p rises slowly with p. According to the ideal gas law, per

mole, $\frac{1}{V} = \frac{p}{RT}$. $1/V$ is thus the concentration of the gas; that is, mol/L.

2.111 As the temperature is increased, the dry ice will sublime, then the pressure will gradually increase. Although the ideal and van der Waals equation do not predict the same pressure at all temperatures, we can calculate the temperature at which they do agree as follows:

From the ideal gas equation:

$$p_{ideal} = \frac{nRT}{V}$$

and from the van der Waals equation:

$$p_{vdW} = \frac{nRT}{V\text{-}nb} - a\frac{n^2}{V^2}$$

If the two equations predict the same pressure, then

$$\frac{nRT}{V} = \frac{nRT}{V\text{-}nb} - a\frac{n^2}{V^2}$$

rearranging to solve for T and simplifying;

$$\frac{nRT}{V\text{-}nb} - \frac{nRT}{V} = a\frac{n^2}{V^2}$$

$$T\left[\frac{nR}{V\text{-}nb} - \frac{nR}{V}\right] = a\frac{n^2}{V^2}$$

$$T = \frac{a\frac{n^2}{V^2}}{\left[\frac{nR}{V\text{-}nb} - \frac{nR}{V}\right]} = \frac{a\frac{n^2}{V^2}}{nR\left[\frac{1}{V\text{-}nb} - \frac{1}{V}\right]} = \frac{\frac{an}{RV^2}}{\left[\frac{1}{V\text{-}nb} - \frac{1}{V}\right]}$$

For CO_2, a = 364.0 kPa L^2 mol^{-2} = 3.64 bar L^2 mol^{-2}; b = 0.0427 L mol^{-1}.

In this problem, V = 50 mL = 0.050 L; n = 36.0 g / 44.0 g mol^{-1} = 0.818 mol

$$T = \frac{\dfrac{3.64\ \text{bar L}^2\text{mol}^2(0.818\ \text{mol})}{0.08314\ \text{L bar K}^{-1}\text{mol}^{-1}(0.050\ \text{L})^2}}{\left[\dfrac{1}{0.050\ \text{L} - 0.818\ \text{mol}(0.0427\ \text{L mol}^{-1})} - \dfrac{1}{0.050\ \text{L}}\right]}$$

$$= \frac{14325\ \text{L K}}{66.35\ \text{L} - 20.00\ \text{L}}$$

$$= 309\ \text{K}$$

At this temperature, the ideal gas equation and the van der Waals equation predict the same pressure, so either can be used to calculate the pressure:

$$p = \frac{nRT}{V} = \frac{0.818\ \text{mol}(0.08314\ \text{L bar K}^{-1}\ \text{mol}^{-1})(309\ \text{K})}{0.050\ \text{L}} = 420\ \text{bar}$$

2.112 Solving this problem requires using both equation 2-3, which relates kinetic energy to molecular speed, and equation 2-4, which relates kinetic energy to temperature. From equation 2-3, we can calculate the kinetic energy of a molecule moving at speed v:

$$E_{\text{kinetic}} = \frac{mv^2}{2}\ \text{(per molecule)}$$

If a collection of atoms has average speed v, then E_{kinetic} is the average kinetic energy of these atoms. For a mole of these atoms, the average kinetic energy is therefore:

$$E_{\text{kinetic}} = N_A \frac{mv^2}{2}\ \text{(per mole)}$$

From equation 2-4, we found that the average molar kinetic energy is related to temperature:

$$\overline{E}_{\text{kinetic}} = \frac{3RT}{2}$$

Equating these two expressions for kinetic energy and solving for the temperature

gives:

$$N_A \frac{mv^2}{2} = \frac{3RT}{2}$$

$$N_A mv^2 = 3RT$$

$$T = \frac{N_A mv^2}{3R} = \frac{Mv^2}{3R}$$

For cesium atoms, M = 0.1329 kg mol^{-1}. Thus,

$$T = \frac{0.1329 \text{ kg mol}^{-1}(0.005 \text{ m s}^{-1})^2}{3(8.314 \text{ J K}^{-1}\text{mol}^{-1})} = 1.3 \times 10^{-7} \text{ K}$$

