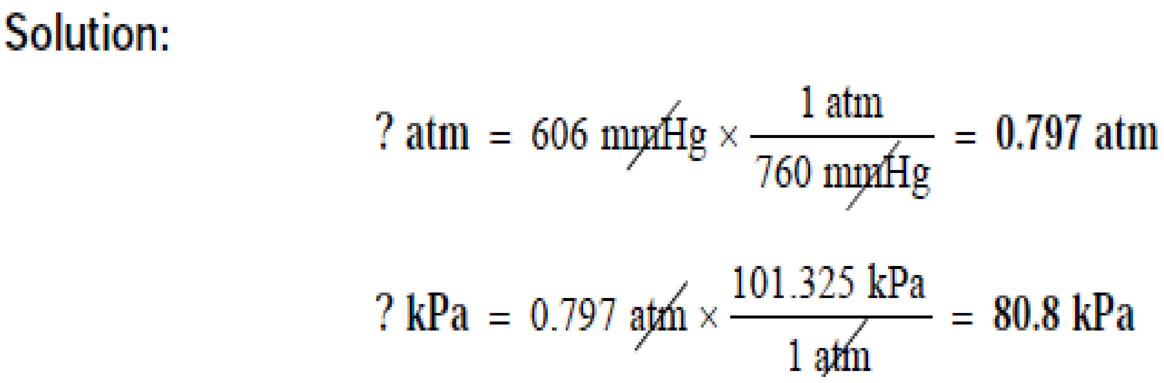
Question & answers
Ch.5

- CHEMISTRY CHANHG, 10TH edition   Raymond Chang

The atmospheric pressure at the summit of Mt. McKinleyis 606 mmHg on a certain day. What is the pressure in atm and in kPa ?



Solution:

$$? \text{ atm} = 606 \cancel{\text{mmHg}} \times \frac{1 \text{ atm}}{760 \cancel{\text{mmHg}}} = \mathbf{0.797 \text{ atm}}$$

$$? \text{ kPa} = 0.797 \cancel{\text{atm}} \times \frac{101.325 \text{ kPa}}{1 \cancel{\text{atm}}} = \mathbf{80.8 \text{ kPa}}$$

5.20

At 46°C a sample of ammonia gas exerts a pressure of 5.3 atm. What is the pressure when the volume of the gas is reduced to one-tenth (0.10) of the original value at the same temperature?

- Temperature and amount of gas do not change in this problem ($T1 = T2$ and $n1 = n2$). Pressure and volume
- change; it is a Boyle's law problem.

- $P_1V_1 = P2V2$
- $V2 = 0.10\ V1$
- $P2 = 5.3\ \text{atm}\ V_1/0.10V1$
- $P2 =$ **53 atm**

5.21: The volume of a gas is 5.80 L, measured at 1.00 atm. What is the pressure of the gas in mmHg if the volume is changed to 9.65 L? (The temperature remains constant.)

P1 = 1.00 atm = 760 mmHg P2 = ?

V1 = 5.80 L V2 = 9.65 L

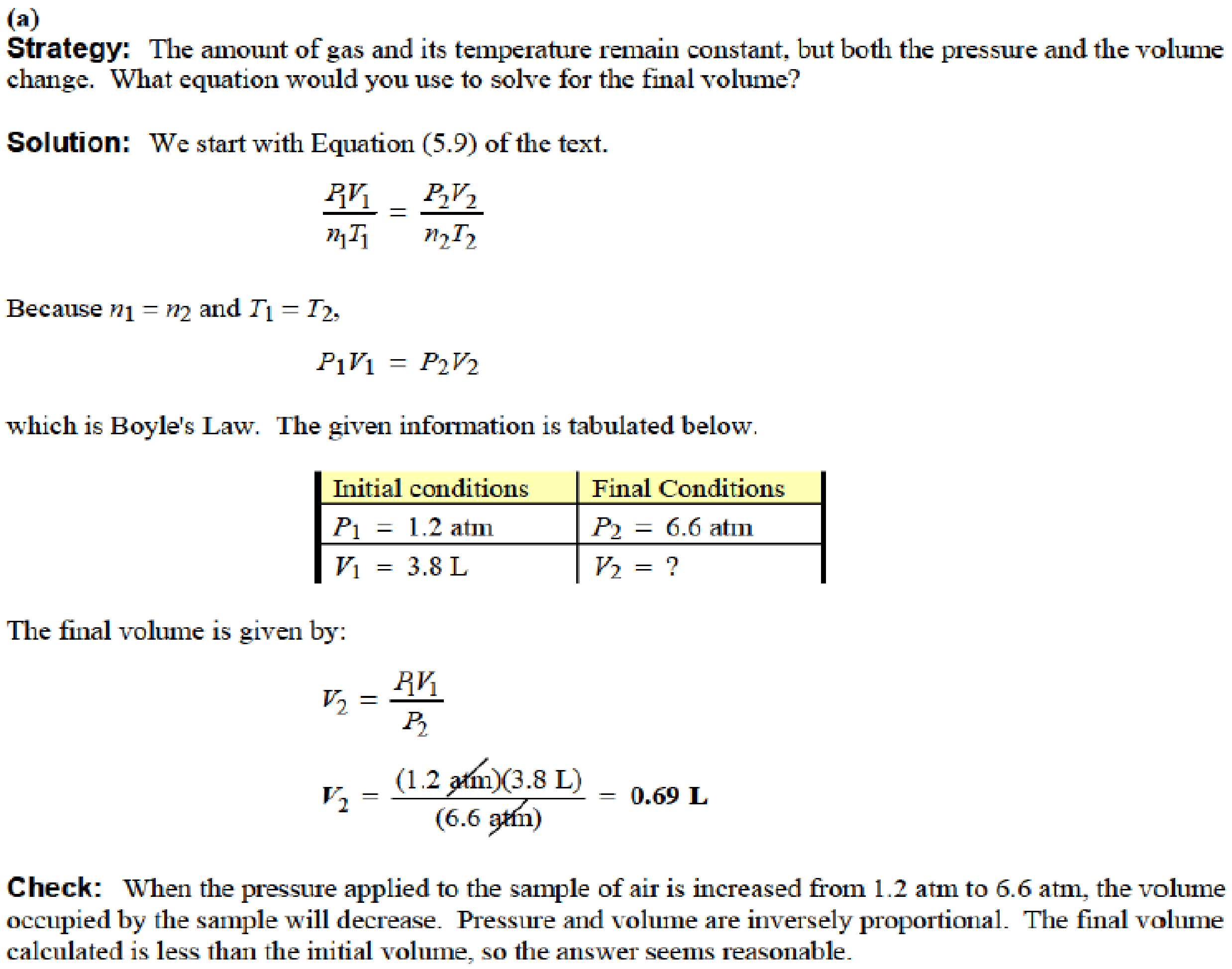
P1V1 = P2V2

P2= P1V1/ V2

(760 mmHg)(5.80 L)/9.65 L

P2= 457 mmHg

5.22: A sample of air occupies 3.8 L when the pressure is 1.2 atm. (a) What volume does it occupy at 6.6 atm? (b) What pressure is required in order to compress it to 0.075 L? (The temperature is kept constant.)



(a)

Strategy: The amount of gas and its temperature remain constant, but both the pressure and the volume change. What equation would you use to solve for the final volume?

Solution: We start with Equation (5.9) of the text.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

Because $n_1 = n_2$ and $T_1 = T_2$,

$$P_1V_1 = P_2V_2$$

which is Boyle's Law. The given information is tabulated below.

Initial conditions	Final Conditions
$P_1 = 1.2$ atm	$P_2 = 6.6$ atm
$V_1 = 3.8$ L	$V_2 = ?$

The final volume is given by:

$$V_2 = \frac{P_1V_1}{P_2}$$

$$V_2 = \frac{(1.2\ \cancel{\text{atm}})(3.8\text{ L})}{(6.6\ \cancel{\text{atm}})} = \mathbf{0.69\ L}$$

Check: When the pressure applied to the sample of air is increased from 1.2 atm to 6.6 atm, the volume occupied by the sample will decrease. Pressure and volume are inversely proportional. The final volume calculated is less than the initial volume, so the answer seems reasonable.

(b)

Strategy: The amount of gas and its temperature remain constant, but both the pressure and the volume change. What equation would you use to solve for the final pressure?

Solution: You should also come up with the equation $P_1V_1 = P_2V_2$ for this problem. The given information is tabulated below.

Initial conditions	Final Conditions
P_1 = 1.2 atm	P_2 = ?
V_1 = 3.8 L	V_2 = 0.075 L

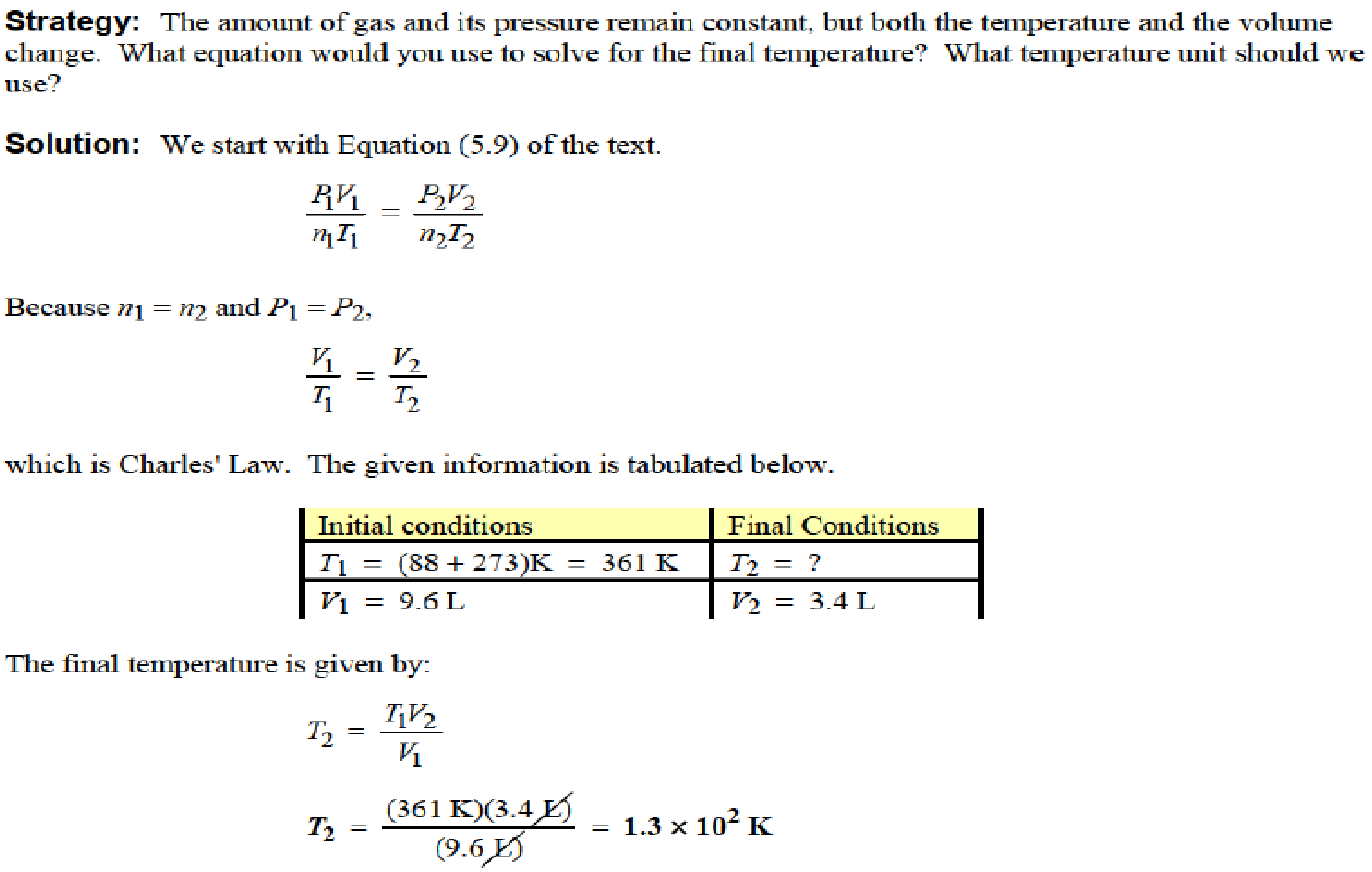
The final pressure is given by:

$$P_2 = \frac{P_1V_1}{V_2}$$

$$\mathbf{P_2} = \frac{(1.2\text{ atm})(3.8\ \cancel{\text{L}})}{(0.075\ \cancel{\text{L}})} = \mathbf{61\ atm}$$

Check: To decrease the volume of the gas fairly dramatically from 3.8 L to 0.075 L, the pressure must be increased substantially. A final pressure of 61 atm seems reasonable.

5.24: Under constant-pressure conditions a sample of hydrogen gas initially at 88°and 9.6 L is cooled until its final volume is 3.4 L. What is its final temperature?



Strategy: The amount of gas and its pressure remain constant, but both the temperature and the volume change. What equation would you use to solve for the final temperature? What temperature unit should we use?

Solution: We start with Equation (5.9) of the text.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

Because $n_1 = n_2$ and $P_1 = P_2$,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

which is Charles' Law. The given information is tabulated below.

Initial conditions	Final Conditions
$T_1 = (88 + 273)\text{K} = 361\text{ K}$	$T_2 = ?$
$V_1 = 9.6\text{ L}$	$V_2 = 3.4\text{ L}$

The final temperature is given by:

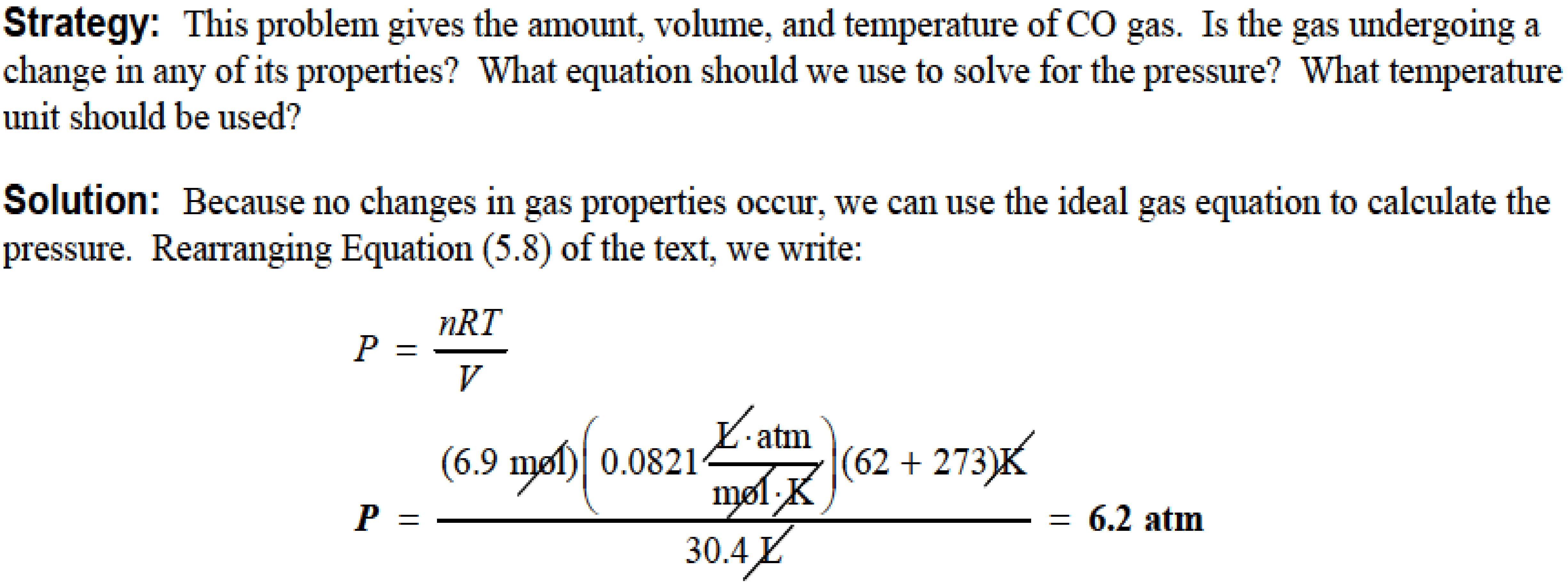
$$T_2 = \frac{T_1V_2}{V_1}$$

$$\boldsymbol{T_2} = \frac{(361\text{ K})(3.4\ \cancel{\text{L}})}{(9.6\ \cancel{\text{L}})} = \mathbf{1.3 \times 10^2\ K}$$

5.31: A sample of nitrogen gas kept in a container of volume 2.3 L and at a temperature of 32°C exerts a pressure of4.7 atm. Calculate the number of moles of gas present.

$$n = \frac{PV}{RT} = \frac{(4.7\ \cancel{\text{atm}})(2.3\ \cancel{\text{L}})}{\left(0.0821 \frac{\cancel{\text{L}} \cdot \cancel{\text{atm}}}{\text{mol} \cdot \cancel{\text{K}}}\right)(273 + 32)\cancel{\text{K}}} = \mathbf{0.43\ mol}$$

5.32:Given that 6.9 moles of carbon monoxide gas are present in a container of volume 30.4 L, what is the pressure of the gas (in atm) if the temperature is 62°C?



Strategy: This problem gives the amount, volume, and temperature of CO gas. Is the gas undergoing a change in any of its properties? What equation should we use to solve for the pressure? What temperature unit should be used?

Solution: Because no changes in gas properties occur, we can use the ideal gas equation to calculate the pressure. Rearranging Equation (5.8) of the text, we write:

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

$$\boldsymbol{P} = \frac{(6.9\ \cancel{\text{mol}})\left(0.0821\,\frac{\cancel{\text{L}}\cdot\text{atm}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}}\right)(62+273)\cancel{\text{K}}}{30.4\ \cancel{\text{L}}} = \textbf{6.2 atm}$$

5.36:The temperature of 2.5 L of a gas initially at STP is raised to 250°C at constant volume. Calculate the final pressure of the gas in atm.

In this problem, the moles of gas and the volume the gas occupies are constant ($V_1 = V_2$ and $n_1 = n_2$). Temperature and pressure change.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

The given information is tabulated below.

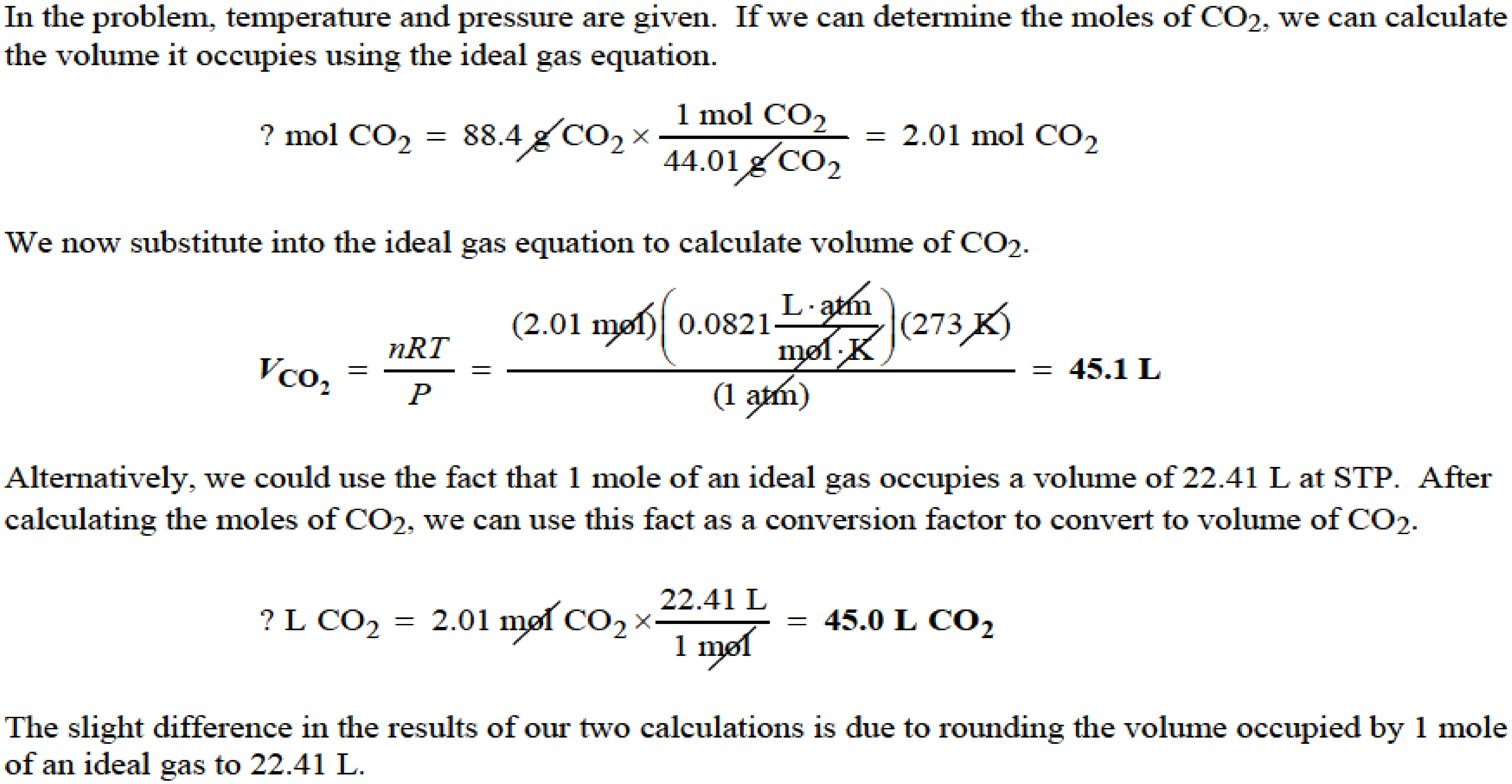
Initial conditions	Final Conditions
$T_1 = 273\text{ K}$	$T_2 = (250 + 273)\text{K} = 523\text{ K}$
$P_1 = 1.0\text{ atm}$	$P_2 = ?$

The final pressure is given by:

$$P_2 = \frac{P_1T_2}{T_1}$$

$$\boldsymbol{P_2} = \frac{(1.0\text{ atm})(523\,\cancel{\text{K}})}{273\,\cancel{\text{K}}} = \mathbf{1.9\ atm}$$

5.40 Calculate its volume (in liters) of 88.4 g of CO_2 at STP.



In the problem, temperature and pressure are given. If we can determine the moles of CO_2, we can calculate the volume it occupies using the ideal gas equation.

$$? \text{ mol } CO_2 = 88.4 \cancel{\text{g}}\, CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \cancel{\text{g}}\, CO_2} = 2.01 \text{ mol } CO_2$$

We now substitute into the ideal gas equation to calculate volume of CO_2.

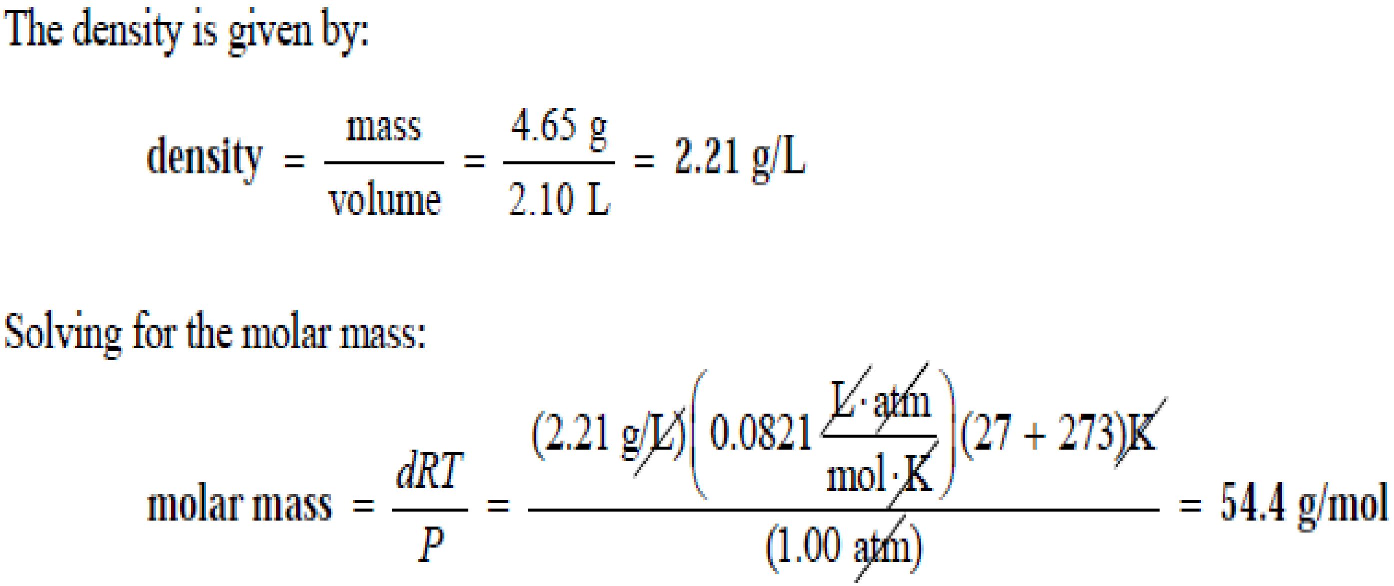
$$V_{CO_2} = \frac{nRT}{P} = \frac{(2.01 \cancel{\text{mol}})\left(0.0821 \frac{\text{L} \cdot \cancel{\text{atm}}}{\cancel{\text{mol}} \cdot \cancel{\text{K}}}\right)(273 \cancel{\text{K}})}{(1 \cancel{\text{atm}})} = \mathbf{45.1\ L}$$

Alternatively, we could use the fact that 1 mole of an ideal gas occupies a volume of 22.41 L at STP. After calculating the moles of CO_2, we can use this fact as a conversion factor to convert to volume of CO_2.

$$? \text{ L } CO_2 = 2.01 \cancel{\text{mol}}\, CO_2 \times \frac{22.41 \text{ L}}{1 \cancel{\text{mol}}} = \mathbf{45.0\ L\ CO_2}$$

The slight difference in the results of our two calculations is due to rounding the volume occupied by 1 mole of an ideal gas to 22.41 L.

5.47:A 2.10-L vessel contains 4.65 g of a gas at 1.00 atm and 27.0°C. (a) Calculate the density of the gas in grams per liter. (b) What is the molar mass of the gas?



The density is given by:

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{4.65\text{ g}}{2.10\text{ L}} = 2.21\text{ g/L}$$

Solving for the molar mass:

$$\text{molar mass} = \frac{dRT}{P} = \frac{(2.21\text{ g/}\cancel{\text{L}})\left(0.0821\dfrac{\cancel{\text{L}}\cdot\cancel{\text{atm}}}{\text{mol}\cdot\cancel{\text{K}}}\right)(27+273)\cancel{\text{K}}}{(1.00\ \cancel{\text{atm}})} = 54.4\text{ g/mol}$$

5.48

Calculate the density of hydrogen bromide (HBr) gas in grams per liter at 733 mmHg and 46°C.

The density can be calculated from the ideal gas equation.

$d = PRT/M$

M = 1.008 g/mol + 79.90 g/mol = 80.91 g/mol

T = 46° + 273° = 319 K

733 mmHg X1 atm / 760 mmHg = 0.964 atm

d =2.98 g/L

Alternatively, we can solve for the density by writing:

density = mass/volume

Assuming that we have 1 mole of HBr, the mass is 80.91 g. The volume of the gas can be calculated using the ideal gas equation.

$V = nRT/P$

(1 mol) x 0.0821 L atm X(319 K)/0.964 atm=

V= 27.2 L

Now, we can calculate the density of HBr gas.

mass 80.91 g

volume 27.2 L

density =2.97 g/L

5.55: A compound of P and F was analyzed as follows: Heating 0.2324 g of the compound in a 378-cm 3 container turned all of it to gas, which had a pressure of 97.3 mmHg at 77°C. Then the gas was mixed with calcium chloride solution, which turned all of the F to 0.2631 g of CaF2 . Determine the molecular formula of the compound

5.55 If you determine the molar mass of the gas, you will be able to determine the molecular formula from the empirical formula. First, let's calculate the molar mass of the compound.
n=PV/RT
(97.3mmHg/760mmHg) x 0.378 L/ 00821 x (77+273K)=.00168 mol

Solving for the molar mass:
mass (in g)/mol = 0.2324 g/ 0.00168 mol
M = **138 g/mol**

To calculate the empirical formula, first we need to find the mass of F in 0.2631 g of CaF2.
02631 gCaF2/78 g/mol CaF2 x 2/1 x 19 =0.128 g

Since the compound only contains P and F, the mass of P in the 0.2324 g sample is:
0.2324 g − 0.1280 g = 0.1044 g P
Now, we can convert masses of P and F to moles of each substance.
? mol P =0.1044 g Px 1 mol P/30.97 g P = 0.003371 mol P

? mol F= 0.1280 g F x 1 mol F/19.00 g= 0.006737 mol F

Thus, we arrive at the formula $P_{0.003371}F_{0.006737}$. Dividing by the smallest number of moles (0.003371 mole) gives the empirical formula PF_2.
To determine the molecular formula, divide the molar mass by the empirical mass.
molar mass 138 g/ 68.97 g= 2
empirical molar mass Hence, the molecular formula is $(PF_2)_2$ or $\mathbf{P_2F_4}$

.63: A mixture of gases contains 0.31 mol CH_4 , 0.25 mol C_2H_6 , and 0.29 mol C_3H_8 . The total pressure is 1.50 atm. Calculate the partial pressures of the gases.

First, we calculate the mole fraction of each component of the mixture. Then, we can calculate the partial pressure of each component using the equation, $P_i = X_i P_T$.

The number of moles of the combined gases is:

$$n = n_{CH_4} + n_{C_2H_6} + n_{C_3H_8} = 0.31 \text{ mol} + 0.25 \text{ mol} + 0.29 \text{ mol} = 0.85 \text{ mol}$$

$$X_{CH_4} = \frac{0.31 \text{ mol}}{0.85 \text{ mol}} = 0.36 \qquad X_{C_2H_6} = \frac{0.25 \text{ mol}}{0.85 \text{ mol}} = 0.29 \qquad X_{C_3H_8} = \frac{0.29 \text{ mol}}{0.85 \text{ mol}} = 0.34$$

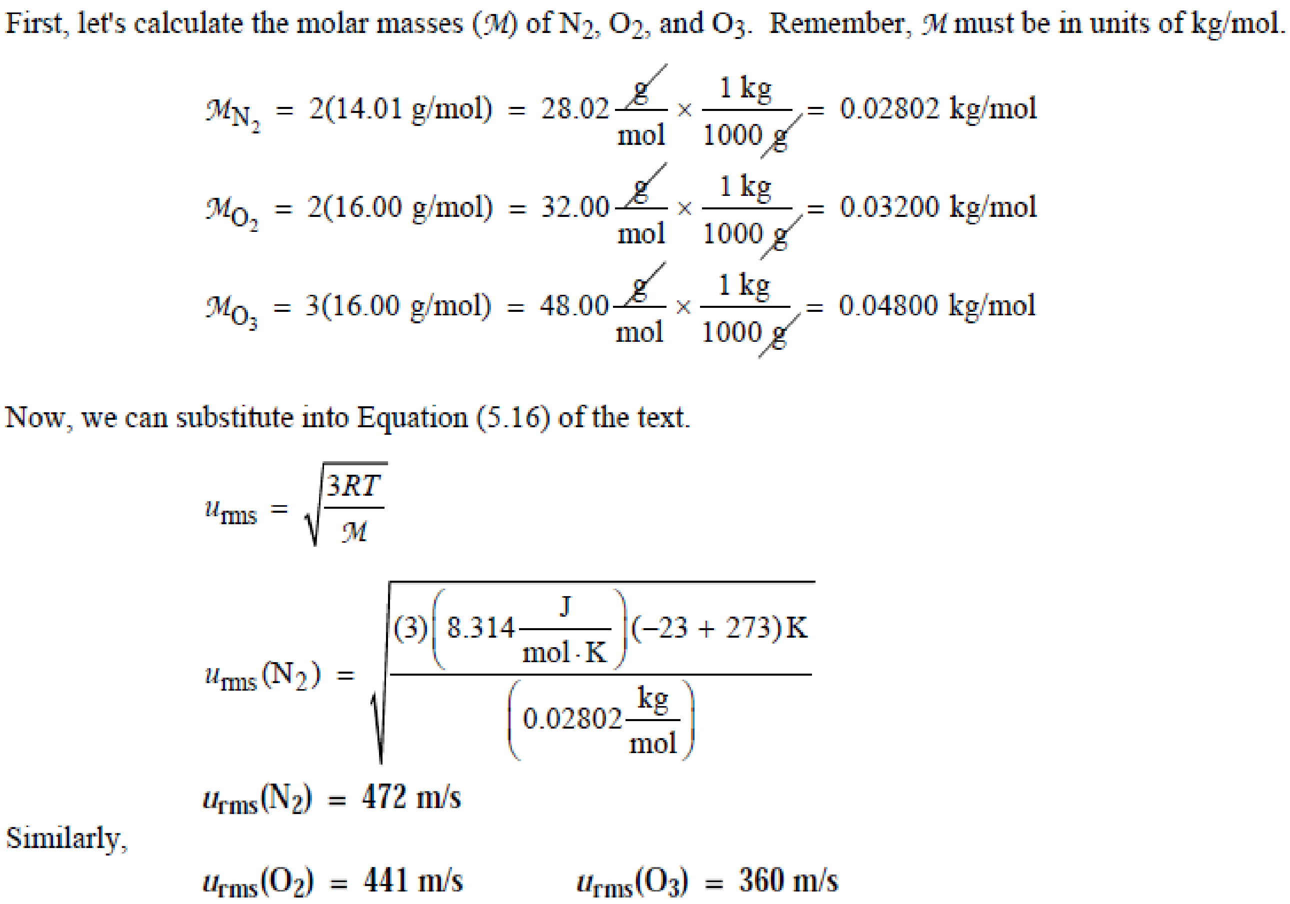
The partial pressures are:

$$\boldsymbol{P_{CH_4}} = X_{CH_4} \times P_{\text{total}} = 0.36 \times 1.50 \text{ atm} = \mathbf{0.54\ atm}$$

$$\boldsymbol{P_{C_2H_6}} = X_{C_2H_6} \times P_{\text{total}} = 0.29 \times 1.50 \text{ atm} = \mathbf{0.44\ atm}$$

$$\boldsymbol{P_{C_3H_8}} = X_{C_3H_8} \times P_{\text{total}} = 0.34 \times 1.50 \text{ atm} = \mathbf{0.51\ atm}$$

5.78:The temperature in the stratosphere is 223°C. Calculate the root-mean-square speeds of N_2, O_2 , and O_3 molecules in this region.



First, let's calculate the molar masses ($\mathcal{M}$) of N_2, O_2, and O_3. Remember, $\mathcal{M}$ must be in units of kg/mol.

$$\mathcal{M}_{N_2} = 2(14.01\text{ g/mol}) = 28.02\frac{\cancel{\text{g}}}{\text{mol}} \times \frac{1\text{ kg}}{1000\ \cancel{\text{g}}} = 0.02802\text{ kg/mol}$$

$$\mathcal{M}_{O_2} = 2(16.00\text{ g/mol}) = 32.00\frac{\cancel{\text{g}}}{\text{mol}} \times \frac{1\text{ kg}}{1000\ \cancel{\text{g}}} = 0.03200\text{ kg/mol}$$

$$\mathcal{M}_{O_3} = 3(16.00\text{ g/mol}) = 48.00\frac{\cancel{\text{g}}}{\text{mol}} \times \frac{1\text{ kg}}{1000\ \cancel{\text{g}}} = 0.04800\text{ kg/mol}$$

Now, we can substitute into Equation (5.16) of the text.

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

$$u_{\text{rms}}(N_2) = \sqrt{\frac{(3)\left(8.314\dfrac{\text{J}}{\text{mol}\cdot\text{K}}\right)(-23 + 273)\text{K}}{\left(0.02802\dfrac{\text{kg}}{\text{mol}}\right)}}$$

$$u_{\text{rms}}(N_2) = 472\text{ m/s}$$

Similarly,

$$u_{\text{rms}}(O_2) = 441\text{ m/s} \qquad u_{\text{rms}}(O_3) = 360\text{ m/s}$$

5.89: Using the data shown in **Table 5.4**, calculate the pressure exerted by 2.50 moles of CO_2 confined in a volume of 5.00 L at 450 K. Compare the pressure with that predicted by the ideal gas equation.

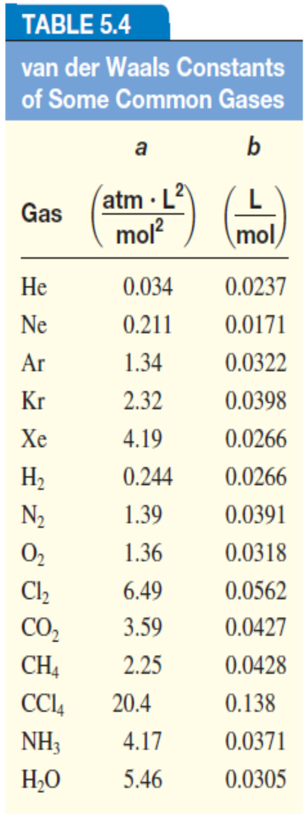
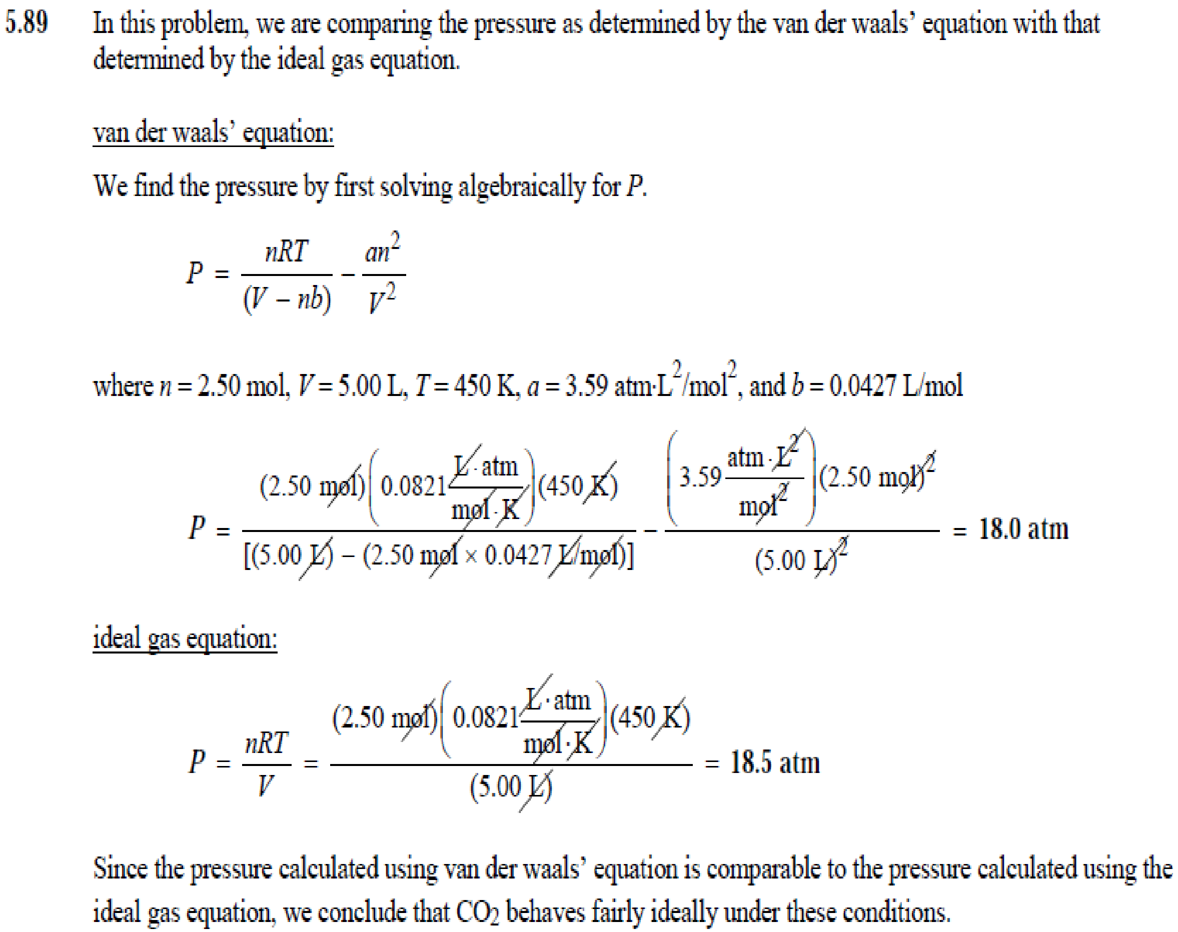


TABLE 5.4

van der Waals Constants of Some Common Gases

Gas	a $\left(\frac{\text{atm}\cdot\text{L}^2}{\text{mol}^2}\right)$	b $\left(\frac{\text{L}}{\text{mol}}\right)$
He	0.034	0.0237
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0266
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
CO_2	3.59	0.0427
CH_4	2.25	0.0428
CCl_4	20.4	0.138
NH_3	4.17	0.0371
H_2O	5.46	0.0305



5.89 In this problem, we are comparing the pressure as determined by the van der waals' equation with that determined by the ideal gas equation.

van der waals' equation:

We find the pressure by first solving algebraically for P.

$$P = \frac{nRT}{(V - nb)} - \frac{an^2}{V^2}$$

where $n = 2.50$ mol, $V = 5.00$ L, $T = 450$ K, $a = 3.59$ atm·L^2/mol^2, and $b = 0.0427$ L/mol

$$P = \frac{(2.50\,\cancel{\text{mol}})\left(0.0821\frac{\cancel{\text{L}}\cdot\text{atm}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}}\right)(450\,\cancel{\text{K}})}{[(5.00\,\cancel{\text{L}}) - (2.50\,\cancel{\text{mol}} \times 0.0427\,\cancel{\text{L}}/\cancel{\text{mol}})]} - \frac{\left(3.59\frac{\text{atm}\cdot\cancel{\text{L}^2}}{\cancel{\text{mol}^2}}\right)(2.50\,\cancel{\text{mol}})^2}{(5.00\,\cancel{\text{L}})^2} = 18.0\text{ atm}$$

ideal gas equation:

$$P = \frac{nRT}{V} = \frac{(2.50\,\cancel{\text{mol}})\left(0.0821\frac{\cancel{\text{L}}\cdot\text{atm}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}}\right)(450\,\cancel{\text{K}})}{(5.00\,\cancel{\text{L}})} = 18.5\text{ atm}$$

Since the pressure calculated using van der waals' equation is comparable to the pressure calculated using the ideal gas equation, we conclude that CO_2 behaves fairly ideally under these conditions.