

- 5.8 The question asks for moles, which can be obtained from P - V - T data using the ideal gas equation: $n = \frac{PV}{RT}$.

Now use the rearranged gas law to determine the number of moles in the sample:

$$n = \frac{PV}{RT} = \frac{(6.47 \times 10^5 \text{ Pa})(5.65 \times 10^{-4} \text{ m}^3)}{(8.314 \frac{\text{J}}{\text{mol K}})(21.7 + 273.15 \text{ K})} = 0.149 \text{ mol.}$$

All conditions except the pressure and volume are fixed, so $P_1V_1 = P_2V_2$ can be used:

$$V_2 = \frac{P_1V_1}{P_2} = \frac{(647 \text{ kPa})(0.565 \text{ L})}{101 \text{ kPa}} = 3.62 \text{ L}$$

- 5.30 The ideal gas equation, $PV = nRT$, can be used to calculate moles using P - V - T data. Molar mass is related to moles through $n = \frac{m}{MM}$. These equations yield two expressions for n ,
- $$n = \frac{PV}{RT} = \frac{m}{MM}, \text{ from which } MM = \frac{mRT}{PV}.$$

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:

$$MM = \frac{(1.65 \text{ g})(8.206 \times 10^{-2} \frac{\text{L atm}}{\text{mol K}})(294.7 \text{ K})}{(1.50 \text{ atm})(0.945 \text{ L})} = 28.1 \text{ g/mol}$$

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 = 4 H. C_2H_4 is the molecular formula.

- 5.86 Although this looks like a problem involving changing conditions, there is enough information given to calculate P directly from the ideal gas equation.

$$n = \frac{m}{MM} = 96.0 \text{ g} \left(\frac{1 \text{ mol}}{32.00 \text{ g}} \right) = 3.00 \text{ mol;}$$

$$V = 3.00 \text{ L;}$$

$$T = 27 + 273 = 300 \text{ K;}$$

$$P = \frac{nRT}{V} = \left(\frac{(3.00 \text{ mol})(8.206 \times 10^{-2} \frac{\text{L atm}}{\text{mol K}})(300 \text{ K})}{3.00 \text{ L}} \right) \left(\frac{760 \text{ torr}}{1 \text{ atm}} \right) = 1.87 \times 10^4 \text{ torr.}$$

5.80 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 $V_f = \frac{P_i V_i}{P_f}$

$$V_f = \frac{(322 \text{ torr})(2.00 \text{ L})}{525 \text{ torr}} = 1.23 \text{ L}$$

(b) Here, pressure, temperature, and volume are changing. So the rearranged ideal gas law is:

$$P_f = \frac{P_i V_i T_f}{V_f T_i}$$

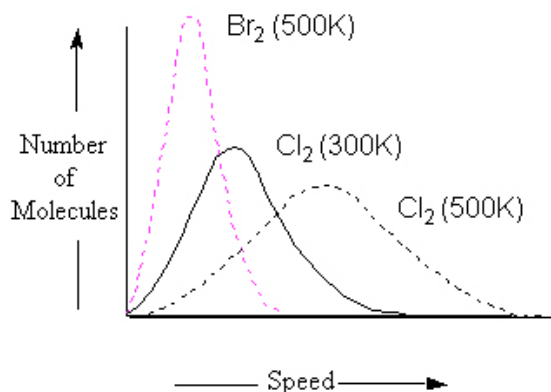
$$P_f = \frac{(322 \text{ torr})(2.00 \text{ L})(50.0 + 273.15)\text{K}}{(1.50 \text{ L})(100.00 + 273.15)\text{K}} = 372 \text{ torr}$$

(c) Here, only pressure and number of moles are changing. So the ideal gas law can be rearranged to:

$$P_f = \frac{P_i n_f}{n_i}$$

$$P_f = \frac{(322 \text{ torr})(1 \text{ mol})}{2 \text{ mol}} = 161 \text{ torr}$$

5.18 Molecular speed is related to temperature and mass by $u_{\text{rms}} = \sqrt{\frac{3RT}{MM}}$ or $u_{\text{mp}} = \sqrt{\frac{2RT}{MM}}$. In either case, the distribution of molecular speeds will have the same general shape, with the position of the peak depending on $\sqrt{\frac{T}{MM}}$. Taking into account the temperatures and molar masses, the graph should look like:



5.22 The root mean square speed per mol of gas is $u_{\text{rms}} = \sqrt{\frac{3RT}{MM}}$ and the average kinetic energy per mol of gas is $E_{\text{kinetic, molar}} = \frac{3}{2}RT$.

(a) Ar: $u_{\text{rms}} = \sqrt{\frac{3(8.314)(400)}{0.039948}} = 5.00 \times 10^2 \text{ m/s per mol}$

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

$$(b) \text{ Xe: } u_{\text{rms}} = \sqrt{\frac{3(8.314)(400)}{0.13129}} = 2.76 \times 10^2 \text{ m/s per mol}$$

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

$$(c) \text{ C}_3\text{H}_8: u_{\text{rms}} = \sqrt{\frac{3(8.314)(600)}{0.04409}} = 5.82 \times 10^2 \text{ m/s};$$

$$E_{\text{kinetic, molar}} = \left(\frac{3}{2}\right)\left(\frac{8.314 \text{ J}}{1 \text{ mol K}}\right)(600\text{K}) = 7.48 \times 10^3 \text{ J/mol}.$$

5.24 Pressure is the cumulative 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, causing the pressure to increase.

(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.

5.74 For these problems we want to use $PV = nRT$.

(a) Since we know that both bulbs have the same volume then

$$\frac{n_{\text{Cl}_2}}{n_{\text{N}_2}} = \frac{P_{\text{Cl}_2} T_{\text{N}_2}}{P_{\text{N}_2} T_{\text{Cl}_2}} = \frac{2 \text{ atm}(298 \text{ K})}{1 \text{ atm}(373 \text{ K})} = 1.6. \text{ Therefore, there are 1.6 more moles of Cl}_2 \text{ than}$$

N_2 , so there are also more molecules of Cl_2 .

(b) To determine mass you need to 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 eqn: $\bar{E}_{\text{kinetic}} = \frac{3RT}{2N_A}$; , the bulb at the higher temperature will have the higher kinetic energy. Again this bulb is the Cl_2 bulb at 100°C compared to the nitrogen bulb at 25°C .

(d) Using the eqn: $u_{\text{rms}} = \left(\frac{3RT}{MM} \right)^{1/2}$ we can determine the average molecular speed for both bulbs.

$$\text{Cl}_2: u_{\text{rms}} = \sqrt{\frac{3(8.314 \text{ J mol}^{-1} \text{ K}^{-1})(373 \text{ K})}{70.90 \times 10^{-3} \text{ kg mol}^{-1}}} = 362 \text{ m/s}$$

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

Therefore, the N₂ bulb has a higher average molecular speed.

5.40 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. The equations needed to solve this problem are:

$$n = \frac{m}{MM}, \quad X_i = \frac{n_i}{n_{\text{tot}}}, \quad p_i = \frac{n_i RT}{V}$$

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

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

$$n(\text{CH}_4) = 1.25 \text{ g} \left(\frac{1 \text{ mol}}{16.04 \text{ g}} \right) = 7.79 \times 10^{-2} \text{ mol};$$

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

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

$$p(\text{Ar}) = \frac{(3.13 \times 10^{-2} \text{ mol})(8.206 \times 10^{-2} \frac{\text{L atm}}{\text{mol K}})(648 \text{ K})}{4.00 \text{ L}} = 0.416 \text{ atm};$$

$$p(\text{CO}) = \frac{(4.46 \times 10^{-2} \text{ mol})(8.206 \times 10^{-2} \frac{\text{L atm}}{\text{mol K}})(648 \text{ K})}{4.00 \text{ L}} = 0.593 \text{ atm};$$

$$p(\text{CH}_4) = \frac{(7.79 \times 10^{-2} \text{ mol})(8.206 \times 10^{-2} \frac{\text{L atm}}{\text{mol K}})(648 \text{ K})}{4.00 \text{ L}} = 1.04 \text{ atm};$$

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

$$P_{\text{tot}} = 0.416 \text{ atm} + 0.593 \text{ atm} + 1.04 \text{ atm} = 2.05 \text{ atm};$$

Determine the mole fraction by dividing the moles of each component by the total number of moles:

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

$$X(\text{Ar}) = \frac{3.13 \times 10^{-2} \text{ mol}}{1.538 \times 10^{-1} \text{ mol}} = 0.204;$$

$$X(\text{CO}) = \frac{4.46 \times 10^{-2} \text{ mol}}{1.538 \times 10^{-1} \text{ mol}} = 0.290;$$

$$X(\text{CH}_4) = \frac{7.79 \times 10^{-2} \text{ mol}}{1.538 \times 10^{-1} \text{ mol}} = 0.507.$$

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

$$X(\text{CO}_2) = (351.5 \text{ ppm})(10^{-6}/\text{ppm}) = 3.515 \times 10^{-4};$$

$$p(\text{CO}_2) = (3.515 \times 10^{-4})(1 \text{ atm}) = 3.515 \times 10^{-4} \text{ atm};$$

$$\frac{n}{V} = \frac{P}{RT} = \frac{3.515 \times 10^{-4} \text{ atm}}{(8.206 \times 10^{-2} \frac{\text{L atm}}{\text{mol K}})(-45 + 273 \text{ K})} = 1.88 \times 10^{-5} \text{ mol/L};$$

$$\#/\text{V} = \left(\frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \right) \left(\frac{1.88 \times 10^{-5} \text{ mol}}{1.0 \text{ L}} \right) = 1.1 \times 10^{19} \text{ molecules/L.}$$

- 5.90 (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:

$$n(\text{N}_2) = \frac{pV}{RT} = \frac{(2.0 \text{ atm})(15 \text{ L})}{(8.206 \times 10^{-2} \frac{\text{L atm}}{\text{mol K}})(300 \text{ K})} = 1.2 \text{ mol};$$

$$n(\text{O}_2) = \frac{pV}{RT} = \frac{(3.0 \text{ atm})(1.5 \text{ L})}{(8.206 \times 10^{-2} \frac{\text{L atm}}{\text{mol K}})(300 \text{ K})} = 0.18 \text{ mol};$$

- (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 $PV = \text{constant}$, thus $P_f V_f = P_i V_i$

$$P_f(\text{N}_2) = \frac{P_i V_i}{V_f} = \frac{(2.0 \text{ atm})(15 \text{ L})}{(15 + 1.5 \text{ L})} = 1.8 \text{ atm};$$

$$P_f(\text{O}_2) = \frac{P_i V_i}{V_f} = \frac{(3.0 \text{ atm})(1.5 \text{ L})}{(15 + 1.5 \text{ L})} = 0.27 \text{ atm};$$

- (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. $\frac{V_{\text{small}}}{V_{\text{total}}} = \frac{1.5 \text{ L}}{1.5 \text{ L} + 15 \text{ L}} = \frac{1}{11}$. There will be 1/11 of them in the smaller container.

- 5.50 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: $2 \text{C}_2\text{H}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{C}_2\text{H}_4\text{O}(\text{g})$

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.

Calculations of initial amounts:

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:

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

$$n_i = \frac{p_i V}{RT} = \frac{(0.500 \text{ atm})(5.00 \times 10^4 \text{ L})}{(8.206 \times 10^{-2} \frac{\text{L atm}}{\text{mol K}})(280 + 273 \text{ K})} = 5.51 \times 10^2 \text{ mol;}$$

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:	$2 \text{C}_2\text{H}_4 (\text{g}) +$	$\text{O}_2(\text{g}) \rightarrow$	$2 \text{C}_2\text{H}_4\text{O} (\text{g})$
Start (10^2 mol)	5.51	5.51	0.00
Change(10^2 mol)	-5.51	-(0.5)(5.51)	+5.51
Final(10^2 mol)	0.00	2.76	5.51

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

$$m(\text{C}_2\text{H}_4\text{O}) = n MM = 5.51 \times 10^2 \text{ mol} \left(\frac{44.05 \text{ g}}{1 \text{ mol}} \right) \left(\frac{10^{-3} \text{ kg}}{1 \text{ g}} \right) = 24.3 \text{ kg.}$$

5.52 The yield in a reaction is the ratio of the actual amount produced to the theoretical amount. The calculation in Problem 5.44 gives the theoretical amount:

$$\text{actual amount} = \text{theoretical amount} \times \frac{\% \text{ yield}}{100\%}$$

$$\text{actual amount} = 24.3 \text{ kg} \left(\frac{65\%}{100\%} \right) = 16 \text{ kg.}$$

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

(a) Begin by converting all values into the units of R

$$P = 759.2 \text{ torr} \left(\frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.9989 \text{ atm}$$

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

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

$$n = \frac{pV}{RT} = \frac{(0.9989 \text{ atm})(0.02296 \text{ L})}{(8.206 \times 10^{-2} \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298.32 \text{ K})} = 9.369 \times 10^{-4} \text{ mol};$$

(b) The balanced reaction is: $2 \text{KClO}_3 \rightarrow 2 \text{KCl} + 3 \text{O}_2$;

$$n(\text{KClO}_3) = 9.369 \times 10^{-4} \text{ mol O}_2 \left(\frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2} \right) = 6.246 \times 10^{-4} \text{ mol};$$

$$(c) m(\text{KClO}_3) = n MM = 6.246 \times 10^{-4} \text{ mol} \left(\frac{122.55 \text{ g}}{1 \text{ mol}} \right) = 7.654 \times 10^{-2} \text{ g};$$

$$\%(\text{KClO}_3) = \frac{m_{\text{KClO}_3}}{m_{\text{total}}} \times 100\% = \frac{7.654 \times 10^{-2} \text{ g}}{0.1054 \text{ g}} \times 100\% = 72.62\%.$$

5.94 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, then use elemental molar masses to find the empirical formula.

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

$$\% \text{ C} = \frac{99.1 \text{ mg}}{125 \text{ mg}} \times 100\% = 79.3 \%;$$

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

$$\% \text{ H} = \frac{7.13 \text{ mg}}{125 \text{ mg}} \times 100\% = 5.70 \%;$$

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

Assume the sample is 100g 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 $\text{C}_7\text{H}_6\text{O}$, with molar mass:

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

Determine the actual molar mass using the ideal gas law. First convert all quantities into the units of R :

$$P = 274 \text{ torr} \left(\frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.361 \text{ atm}$$

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

$$T = 150 + 273 \text{ K} = 423 \text{ K}$$

- 5.68 Assume that "warm moist" means that the air is saturated with water vapor at the collection temperature, 30 °C. From Table 5-4, the partial pressure of water at this temperature is 31.824 torr. Addition of a drying agent removes all of this water vapor:
(a) $P_{\text{final}} = 756 \text{ torr} - 31.824 \text{ torr} = 724 \text{ torr}$;

$$(b) X(\text{H}_2\text{O}) = \frac{P_{\text{H}_2\text{O}}}{P_{\text{total}}} = \left(\frac{31.824 \text{ torr}}{756 \text{ torr}} \right) = 4.21 \times 10^{-2};$$

Use the ideal gas equation to determine the mass density of water vapor:

$$pV = nRT = \frac{mRT}{MM} \text{ can be rearranged to give } \frac{m}{V} = \frac{p MM}{RT};$$

$$p = 31.824 \text{ torr} \left(\frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.041874 \text{ atm}$$

$$\frac{m}{V} = \frac{p MM}{RT} = \frac{(0.041874 \text{ atm})(18.01 \text{ g/mol})}{(8.206 \times 10^{-2} \frac{\text{L atm}}{\text{mol K}})(30 + 273 \text{ K})} = 3.03 \times 10^{-2} \text{ g/L.}$$