

27. a. $4.8 \text{ atm} \times \frac{760 \text{ mm Hg}}{\text{atm}} = 3.6 \times 10^3 \text{ mm Hg}$ b. $3.6 \times 10^3 \text{ mm Hg} \times \frac{1 \text{ torr}}{\text{mm Hg}}$
 $= 3.6 \times 10^3 \text{ torr}$

c. $4.8 \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 4.9 \times 10^5 \text{ Pa}$ d. $4.8 \text{ atm} \times \frac{14.7 \text{ psi}}{\text{atm}} = 71 \text{ psi}$

36. As NO_2 is converted completely into N_2O_4 , the moles of gas present will decrease by one-half (from the 2:1 mol ratio in the balanced equation). Using Avogadro's law,

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}, \quad V_2 = V_1 \times \frac{n_2}{n_1} = 25.0 \text{ mL} \times \frac{1}{2} = 12.5 \text{ mL}$$

$\text{N}_2\text{O}_4(\text{g})$ will occupy one-half the original volume of $\text{NO}_2(\text{g})$. This is expected since the mol of gas present decrease by one-half when NO_2 is converted into N_2O_4 .

40. $\frac{PV}{nT} = R$; For a gas at two conditions:

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}; \text{ Since } n \text{ and } V \text{ are constant: } \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$T_2 = \frac{P_2 T_1}{P_1} = \frac{2500 \text{ torr} \times 294.2 \text{ K}}{758 \text{ torr}} = 970 \text{ K} = 7.0 \times 10^2 \text{ EC}$$

54. Since the solution is 50.0% H_2O_2 by mass, the mass of H_2O_2 decomposed is $125/2 = 62.5 \text{ g}$.

$$62.5 \text{ g H}_2\text{O}_2 \times \frac{1 \text{ mol H}_2\text{O}_2}{34.02 \text{ g H}_2\text{O}_2} \times \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}_2} = 0.919 \text{ mol O}_2$$

$$V = \frac{nRT}{P} = \frac{0.919 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times 300. \text{ K}}{746 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}} = 23.0 \text{ L O}_2$$

58. For ammonia (in one minute):

$$n_{\text{NH}_3} = \frac{PV}{RT} = \frac{90. \text{ atm} \times 500. \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times 496 \text{ K}} = 1.1 \times 10^3 \text{ mol NH}_3$$

NH_3 flows into the reactor at a rate of $1.1 \times 10^3 \text{ mol/min}$.

For CO_2 (in one minute):

$$n_{\text{CO}_2} = \frac{PV}{RT} = \frac{45 \text{ atm} \times 600. \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times 496 \text{ K}} = 6.6 \times 10^2 \text{ mol CO}_2$$

CO₂ flows into the reactor at 6.6×10^2 mol/min.

To react completely with 1.1×10^3 mol NH₃/min, we need:

$$\frac{1.1 \times 10^3 \text{ mol NH}_3}{\text{min}} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol NH}_3} = 5.5 \times 10^2 \text{ mol CO}_2/\text{min}$$

Since 660 mol CO₂/min are present, ammonia is the limiting reagent.

$$\frac{1.1 \times 10^3 \text{ mol NH}_3}{\text{min}} \times \frac{1 \text{ mol urea}}{2 \text{ mol urea}} \times \frac{60.06 \text{ g urea}}{\text{mol urea}} = 3.3 \times 10^4 \text{ g urea/min}$$

64. $d = P \times (\text{molar mass})/RT$; We need to determine the average molar mass of air. We get this by using the mol fraction information to determine the weighted value for the molar mass. If we have 1.000 mol of air:

$$\begin{aligned} \text{average molar mass} &= 0.78 \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{\text{mol N}_2} + 0.21 \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{\text{mol O}_2} + \\ &\quad 0.010 \text{ mol Ar} \times \frac{39.95 \text{ g Ar}}{\text{mol Ar}} = 28.98 = 29 \text{ g} \end{aligned}$$

$$d_{\text{air}} = \frac{1.00 \text{ atm} \times 29 \text{ g/mol}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times 273 \text{ K}} = 1.3 \text{ g/L}$$

67. Use the relationship $P_1V_1 = P_2V_2$ for each gas, since T and n for each gas are constant.

$$\text{For H}_2: P_2 = \frac{P_1V_1}{V_2} = 475 \text{ torr} \times \frac{2.00 \text{ L}}{3.00 \text{ L}} = 317 \text{ torr}$$

$$\text{For N}_2: P_2 = 0.200 \text{ atm} \times \frac{1.00 \text{ L}}{3.00 \text{ L}} = 0.0667 \text{ atm}; 0.0667 \text{ atm} \times \frac{760 \text{ torr}}{\text{atm}} = 50.7 \text{ torr}$$

$$P_{\text{total}} = P_{\text{H}_2} + P_{\text{N}_2} = 317 + 50.7 = 368 \text{ torr}$$

70. If we had 100.0 g of the gas, we would have 50.0 g He and 50.0 g Xe.

$$\chi_{\text{He}} = \frac{n_{\text{He}}}{n_{\text{He}} + n_{\text{Xe}}} = \frac{\frac{50.0 \text{ g}}{4.003 \text{ g/mol}}}{\frac{50.0 \text{ g}}{4.003 \text{ g/mol}} + \frac{50.0 \text{ g}}{131.3 \text{ g/mol}}} = \frac{12.5 \text{ mol He}}{12.5 \text{ mol He} + 0.381 \text{ mol Xe}} = 0.970$$

$$P_{\text{He}} = \chi_{\text{He}} P_{\text{total}} = 0.970 \times 600. \text{ torr} = 582 \text{ torr}; P_{\text{Xe}} = 600. - 582 = 18 \text{ torr}$$

71. $P_{\text{TOT}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}}, 1.032 \text{ atm} = P_{\text{H}_2} + 32 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}, P_{\text{H}_2} = 1.032 - 0.042 = 0.990 \text{ atm}$

$$n_{\text{H}_2} = \frac{P_{\text{H}_2} V}{RT} = \frac{0.990 \text{ atm} \times 0.240 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times 303 \text{ K}} = 9.56 \times 10^{-3} \text{ mol H}_2$$

$$9.56 \times 10^{-3} \text{ mol H}_2 \times \frac{1 \text{ mol Zn}}{\text{mol H}_2} \times \frac{65.38 \text{ g Zn}}{\text{mol Zn}} = 0.625 \text{ g Zn}$$

72. To calculate the volume of gas, we can use P_{total} and n_{total} ($V = n_{\text{tot}}RT/P_{\text{tot}}$) or we can use P_{He} and n_{He} ($V = n_{\text{He}}RT/P_{\text{He}}$). Since $n_{\text{H}_2\text{O}}$ is unknown, we will use P_{He} and n_{He} .

$$P_{\text{He}} + P_{\text{H}_2\text{O}} = 1.00 \text{ atm} = 760. \text{ torr} = P_{\text{He}} + 23.8 \text{ torr}, P_{\text{He}} = 736 \text{ torr}$$

$$n_{\text{He}} = 0.586 \text{ g} \times \frac{1 \text{ mol}}{4.003 \text{ g}} = 0.146 \text{ mol He}$$

$$V = \frac{n_{\text{He}}RT}{P_{\text{He}}} = \frac{0.146 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times 298 \text{ K}}{736 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}} = 3.69 \text{ L}$$

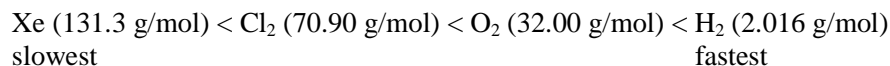
77. $\text{KE}_{\text{avg}} = 3/2 RT$; The average kinetic energy depends only on temperature. At each temperature, CH_4 and N_2 will have the same average KE. For energy units of joules (J), use $R = 8.3145 \text{ J/mol}\cdot\text{K}$. To determine average KE per molecule, divide by Avogadro's number, $6.022 \times 10^{23} \text{ molecules/mol}$.

$$\text{at } 273 \text{ K: } \text{KE}_{\text{avg}} = \frac{3}{2} \times \frac{8.3145 \text{ J}}{\text{mol K}} \times 273 \text{ K} = 3.40 \times 10^3 \text{ J/mol} = 5.65 \times 10^{-21} \text{ J/molecule}$$

$$\text{at } 546 \text{ K: } \text{KE}_{\text{avg}} = \frac{3}{2} \times \frac{8.3145 \text{ J}}{\text{mol K}} \times 546 \text{ K} = 6.81 \times 10^3 \text{ J/mol} = 1.13 \times 10^{-20} \text{ J/molecule}$$

84. a. All the gases have the same average kinetic energy since they are all at the same temperature.

- b. At constant T, the lighter the gas molecule, the faster the average velocity.



- c. At constant T, the lighter H₂ molecules have a faster average velocity than the heavier O₂ molecules. As temperature increases, the average velocity of the gas molecules increases. Separate samples of H₂ and O₂ can only have the same average velocities if the temperature of the O₂ sample is greater than the temperature of the H₂ sample.

88. $\frac{\text{Rate}_1}{\text{Rate}_2} = \left(\frac{M_2}{M_1} \right)^{1/2}$, where M = molar mass; Let Gas (1) = He, Gas (2) = Cl₂:

$$\frac{\frac{1.0 \text{ L}}{4.5 \text{ min}}}{\frac{1.0 \text{ L}}{t}} = \left(\frac{70.90}{4.003} \right)^{1/2}, \quad \frac{t}{4.5 \text{ min}} = 4.209, \quad t = 19 \text{ min}$$

89. a. $P = \frac{nRT}{V} = \frac{0.5000 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times (25.0 + 273.2) \text{ K}}{1.0000 \text{ L}} = 12.24 \text{ atm}$

b. $\left[P + a \left(\frac{n}{V} \right)^2 \right] \times (V - nb) = nRT$; For N₂: $a = 1.39 \text{ atm L}^2/\text{mol}^2$ and $b = 0.0391 \text{ L/mol}$

$$\left[P + 1.39 \left(\frac{0.5000}{1.0000} \right)^2 \text{ atm} \right] \times (1.0000 \text{ L} - 0.5000 \times 0.0391 \text{ L}) = 12.24 \text{ L atm}$$

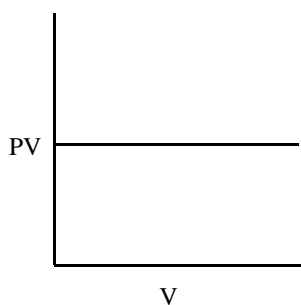
$$(P + 0.348 \text{ atm}) \times (0.9805 \text{ L}) = 12.24 \text{ L atm}$$

$$P = \frac{12.24 \text{ L atm}}{0.9805 \text{ L}} - 0.348 \text{ atm} = 12.48 - 0.348 = 12.13 \text{ atm}$$

- c. The ideal gas law is high by 0.11 atm or $\frac{0.11}{12.13} \times 100 = 0.91\%$.

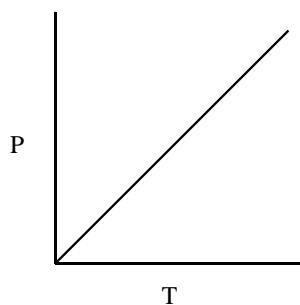
95. a. $PV = nRT$

$PV = \text{constant}$



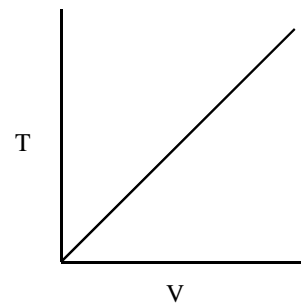
b. $P = \left(\frac{nR}{V} \right) \times T$

$P = \text{constant} \times T$



c. $T = \left(\frac{P}{nR} \right) \times V$

$T = \text{constant} \times V$



d. $PV = nRT$

$PV = \text{constant}$

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

$P = \text{constant} \times \frac{1}{V}$

f. $\frac{PV}{T} = nR$

$\frac{PV}{T} = \text{constant}$

106. For NH_3 : $P_2 = \frac{P_1 V_1}{V_2} = 0.500 \text{ atm} \times \frac{2.00 \text{ L}}{3.00 \text{ L}} = 0.333 \text{ atm}$

For O_2 : $P_2 = \frac{P_1 V_1}{V_2} = 1.50 \text{ atm} \times \frac{1.00 \text{ L}}{3.00 \text{ L}} = 0.500 \text{ atm}$

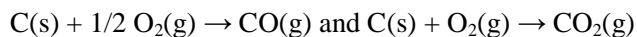
After the stopcock is opened, V and T will be constant, so $P \propto n$. The balanced equation requires:

$$\frac{n_{\text{O}_2}}{n_{\text{NH}_3}} = \frac{P_{\text{O}_2}}{P_{\text{NH}_3}} = \frac{5}{4} = 1.25$$

The actual ratio present is: $\frac{P_{\text{O}_2}}{P_{\text{NH}_3}} = \frac{0.500 \text{ atm}}{0.333 \text{ atm}} = 1.50$

The actual ratio is larger than the required ratio, so NH_3 in the denominator is limiting. Because equal moles of NO will be produced as NH_3 that reacts, the partial pressure of NO produced is 0.333 atm (the same as P_{NH_3} reacted).

120. The reactions are:



$$PV = nRT, P = n \left(\frac{RT}{V} \right) = n (\text{constant})$$

Since the pressure has increased by 17.0%, the number of moles of gas has also increased by 17.0%.

$$n_{\text{final}} = 1.170 n_{\text{initial}} = 1.170 (5.00) = 5.85 \text{ mol gas} = n_{\text{O}_2} + n_{\text{CO}} + n_{\text{CO}_2}$$

$n_{\text{CO}} + n_{\text{CO}_2} = 5.00$ (balancing moles of C). Solving by simultaneous equations:

$$\begin{array}{rcl} n_{\text{O}_2} + n_{\text{CO}} + n_{\text{CO}_2} & = & 5.85 \\ (n_{\text{CO}} + n_{\text{CO}_2}) & = & 5.00 \\ \hline n_{\text{O}_2} & = & 0.85 \end{array}$$

If all C was converted to CO_2 , no O_2 would be left. If all C was converted to CO, we would get 5 mol CO and 2.5 mol excess O_2 in the reaction mixture. In the final mixture, mol of CO equals twice the mol O_2 present ($n_{\text{CO}} = 2 n_{\text{O}_2}$).

$$n_{\text{CO}} = 2 n_{\text{O}_2} = 1.70 \text{ mol CO}; 1.70 + n_{\text{CO}_2} = 5.00, n_{\text{CO}_2} = 3.30 \text{ mol CO}_2$$

$$\chi_{\text{CO}} = \frac{1.70}{5.85} = 0.291; \chi_{\text{CO}_2} = \frac{3.30}{5.85} = 0.564; \chi_{\text{O}_2} = \frac{0.85}{5.85} = 0.145$$

130. The partial pressures can be determined by using the mole fractions.

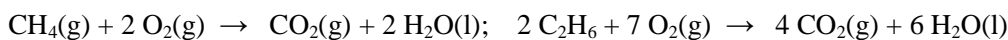
$$P_{\text{methane}} = P_{\text{tot}} \times \chi_{\text{methane}} = 1.44 \text{ atm} \times 0.915 = 1.32 \text{ atm}; P_{\text{ethane}} = 1.44 - 1.32 = 0.12 \text{ atm}$$

Determining the number of mol of natural gas combusted:

$$n_{\text{natural gas}} = \frac{PV}{RT} = \frac{1.44 \text{ atm} \times 15.00 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times 293 \text{ K}} = 0.898 \text{ mol natural gas}$$

$$n_{\text{methane}} = n_{\text{natural gas}} \times \chi_{\text{methane}} = 0.898 \text{ mol} \times 0.915 = 0.822 \text{ mol methane}$$

$$n_{\text{ethane}} = 0.898 - 0.822 = 0.076 \text{ mol ethane}$$



$$0.822 \text{ mol CH}_4 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CH}_4} \times \frac{18.02 \text{ g H}_2\text{O}}{\text{mol H}_2\text{O}} = 29.6 \text{ g H}_2\text{O}$$

$$0.076 \text{ mol C}_2\text{H}_6 \times \frac{6 \text{ mol H}_2\text{O}}{2 \text{ mol C}_2\text{H}_6} \times \frac{18.02 \text{ g H}_2\text{O}}{\text{mol H}_2\text{O}} = 4.1 \text{ g H}_2\text{O}$$

The total mass of H₂O produced = 29.6 g + 4.1 g = 33.7 g H₂O

