

Chapter 6
Solutions to Selected Integrative and Advanced Exercises

54. (M)

(a) In this case, H₂ is the limiting reagent.

$$\text{Volume NH}_3 = 313 \text{ L H}_2 \times \frac{2 \text{ L NH}_3}{3 \text{ L H}_2} = 209 \text{ L NH}_3$$

(b) Moles of NH₃ (@ 315 °C and 5.25 atm)

$$= \frac{5.25 \text{ atm} \times 313 \text{ L}}{0.08206 \frac{\text{L atm}}{\text{mol K}} \times (315 + 273)\text{K}} = 3.41 \times 10^1 \text{ mol H}_2 \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} = 2.27 \times 10^1 \text{ mol NH}_3$$

$$V(@25^\circ\text{C}, 727\text{mmHg}) = \frac{2.27 \times 10^1 \text{ mol NH}_3 \times 0.08206 \frac{\text{L atm}}{\text{mol K}} \times 298 \text{ K}}{727 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 5.80 \times 10^2 \text{ L NH}_3$$

65. (M) In this problem, you don't need to explicitly solve for moles of gas, since you are looking at the relationship between pressure and volume.

$$\text{mol O}_2 = \frac{PV}{RT} = \frac{(4.0 \text{ atm})(1.0 \text{ L})}{RT} = \frac{4.0}{RT}$$

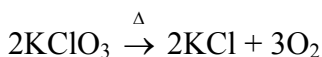
$$\text{mol N}_2 = \frac{PV}{RT} = \frac{(2.0 \text{ atm})(2.0 \text{ L})}{RT} = \frac{4.0}{RT}$$

$$\text{total mol. of gas} = 8.0/RT$$

Therefore,

$$P = \frac{nRT}{V} = \frac{8.0}{RT} \cdot \frac{RT}{2.0} = 4.0 \text{ atm}$$

72. (M) The work for this problem is nearly identical to the above problem.



The pressure of O₂ is determined by subtracting the known vapor pressure of water at the given temperature from the measured total pressure.

$$P_{\text{O}_2} = P_{\text{TOT}} - P_{\text{H}_2\text{O}} = 323 \text{ torr} - 25.22 \text{ torr} = 298 \text{ torr}$$

$$P_{\text{atm}} = 298 \text{ mmHg} / 760 \text{ mmHg} = 0.392 \text{ atm}$$

$$\text{mol O}_2 = \frac{PV}{RT} = \frac{(0.392 \text{ atm})(0.229 \text{ L})}{(0.08206 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(299 \text{ K})} = 0.003657 \text{ mol}$$

Mass of KClO₃ is then determined as follows:

$$0.003657 \text{ mol O}_2 \times \frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2} \times \frac{122.54 \text{ g}}{1 \text{ mol KClO}_3} = 0.299 \text{ g KClO}_3$$

$$\% \text{KClO}_3 = \frac{0.299 \text{ g}}{0.415 \text{ g}} \times 100 = 72\%$$

95. (M) First, calculate the number of moles of gas:

$$n = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(6.30 \text{ L})}{(0.082058 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(323 \text{ K})} = 0.2377 \text{ mol}$$

Molecular mass (MM) then can be calculated:

$$\text{MM} = (10.00 \text{ g}) / (0.2377 \text{ mol}) = 42.069 \text{ g/mol}$$

Now we must determine the mole ratio of C to H:

$$85.6 \text{ g C} \times (1 \text{ mol C} / 12.01 \text{ g C}) = 7.13 \text{ mol}$$

$$14.4 \text{ g H} \times (1 \text{ mol H} / 1.01 \text{ g H}) = 14.3 \text{ mol}$$

Now, divide both by the smallest number:

$$\text{C:C mole ratio} = 1$$

$$\text{H:C mole ratio} = 2$$

Therefore, the empirical formula is CH_2 , with a formula unit molar mass of $\sim 14.03 \text{ g/mol}$. To determine molecular formula, divide MM by formula unit MM:

$$42.069 / 14.03 = 3.00. \text{ Therefore, the molecular formula is } \text{C}_3\text{H}_6.$$

96. (M) First we determine the molar mass of the hydrocarbon.

$$M = \frac{mRT}{PV} = \frac{0.7178 \text{ g} \times 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} (65.0 + 273.2) \text{ K}}{\left(99.2 \text{ kPa} \times \frac{0.00987 \text{ atm}}{1 \text{ kPa}}\right) \times 0.3907 \text{ L}} = 52.0 \text{ g/mol}$$

Now determine the empirical formula. A hydrocarbon contains just hydrogen and carbon.

$$\text{amount C} = 2.4267 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.05514 \text{ mol C} \div 0.05514 \rightarrow 1.000 \text{ mol C}$$

$$\text{amount H} = 0.4967 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.05513 \text{ mol H} \div 0.05514 \rightarrow 1.000 \text{ mol H}$$

The empirical formula is CH. This gives an empirical molar mass of 13 g/mol , almost precisely one-fourth of the experimental molar mass. The molecular formula is therefore C_4H_4 .

98. (M) First, let's convert the given units to those easier used:

$$P = 101 \text{ kPa} \times (1 \text{ barr} / 101 \text{ kPa}) \times (1 \text{ atm} / 1.01 \text{ barr}) = 0.9901 \text{ atm}$$

$$T = 819 \text{ }^\circ\text{C} + 273 \text{ K} = 1092 \text{ K}$$

$$\text{mol NH}_4\text{NO}_2 = 128 \text{ g NH}_4\text{NO}_2 \times \frac{1 \text{ mol NH}_4\text{NO}_2}{64.052 \text{ g}} = 1.998 \text{ mol}$$

$$\text{mol gas} = 1.998 \text{ mol NH}_4\text{NO}_2 \times \frac{3 \text{ mol gas}}{1 \text{ mol NH}_4\text{NO}_2} = 5.994$$

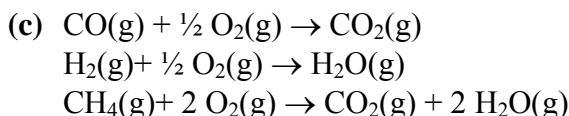
$$V = \frac{nRT}{P} = \frac{(5.994 \text{ mol})(0.082058 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(1092 \text{ K})}{0.9901 \text{ atm}} = 542 \text{ L}$$

107. (D)

$$(a) M_{av} = \frac{8.0}{100} \left(44.01 \frac{\text{g}}{\text{mol}} \right) + \frac{23.2}{100} \left(28.01 \frac{\text{g}}{\text{mol}} \right) + \frac{17.7}{100} \left(2.016 \frac{\text{g}}{\text{mol}} \right) + \frac{1.1}{100} \left(16.043 \frac{\text{g}}{\text{mol}} \right) + \frac{50.5}{100} \left(28.01 \frac{\text{g}}{\text{mol}} \right)$$

$$M_{av} = 24.56 \text{ g/mol} \quad \text{density} = \frac{PM}{RT} = \frac{\left(763 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) 24.56 \frac{\text{g}}{\text{mol}}}{\left(0.08206 \frac{\text{L atm}}{\text{K mol}} \right) 296 \text{ K}} = 1.015 \frac{\text{g}}{\text{L}}$$

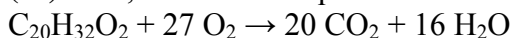
$$(b) P_{co} = P_{total} \times \frac{V\%}{100\%} = 763 \text{ mmHg} \times \frac{23.2\%}{100\%} = 177 \text{ mmHg or } 0.233 \text{ atm}$$



Use the fact that volume is directly proportional to moles when the pressure and temperature are constant. 1000 L of producer gas contains:

232 L CO	which requires	116 L O ₂
177 L H ₂	which requires	88.5 L O ₂
11 L CH ₄	which requires	<u>22 L O₂</u>
		226.5 L O ₂ (Note: air is 20.95 % O ₂ by volume)

$$\text{Thus, the reaction requires } \frac{226.5 \text{ L}}{0.2095} = 1.08 \times 10^3 \text{ L.}$$

108. (M) First, balance the equation:

$$\text{mol O}_2 \text{ needed: } 2000 \text{ g C}_{20}\text{H}_{32}\text{O}_2 \times \frac{1 \text{ mol C}_{20}\text{H}_{32}\text{O}_2}{304.52 \text{ g C}_{20}\text{H}_{32}\text{O}_2} \times \frac{27 \text{ mol O}_2}{1 \text{ mol C}_{20}\text{H}_{32}\text{O}_2}$$

$$= 177.33 \text{ mol O}_2$$

Using the ideal gas law, we can determine the volume of 177.33 mol of O₂:

$$\text{vol O}_2 = \frac{nRT}{P} = \frac{(177.33 \text{ mol})(0.082058 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(298 \text{ K})}{1.00 \text{ atm}} = 4336.30 \text{ L O}_2$$

To determine the volume of air needed, we note that O₂ represents 20.9% of air by volume: $x(0.205) = 4336.30 \text{ L}$. Solving for x gives 20698 L, or $2.070 \times 10^4 \text{ L}$.

113. (M) First, determine the moles of Cl₂ and NaClO. Then, determine the limiting reagent

$$\text{mol Cl}_2 = \frac{(4.66 \text{ atm})(1.0 \text{ L})}{(0.082058 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1})(283 \text{ K})} = 0.2007 \text{ mol}$$

$$\text{mol NaClO} = (0.750 \text{ L})(2.00 \text{ M}) = 1.50 \text{ mol}$$

$$0.2007 \text{ mol Cl}_2 \times \frac{4 \text{ mol NaClO}}{1 \text{ mol Cl}_2} = 0.8028 \text{ mol NaClO needed}$$

Therefore, NaClO is the excess reagent.

Now, we must calculate the theoretical yield of ClO₂.

$$0.2007 \text{ mol Cl}_2 \times \frac{2 \text{ mol ClO}_2}{1 \text{ mol Cl}_2} \times \frac{67.45 \text{ g ClO}_2}{1 \text{ mol ClO}_2} = 27.07 \text{ g ClO}_2$$

$$\% \text{yield} = \frac{25.9}{27.07} \times 100 = 95.7\%$$

119. (D) Potential energy of an object is highest when the kinetic energy of the object is zero and the object has attained its maximum height. Therefore, we must determine the kinetic energy. But first, we have to determine the velocity of the N_2 molecule.

$$\overline{u_{\text{rms}}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3(8.3145 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1})(300 \text{ K})}{28.00 \times 10^{-3} \text{ kg}}} = 517 \text{ m/s}$$

$$\overline{e_k} = \frac{1}{2} m \overline{u_{\text{rms}}}^2 = \frac{1}{2} (28.00 \times 10^{-3} \text{ kg})(517 \text{ m/s})^2 = 3742 \text{ J}$$

$$\overline{e_k} = \overline{e_p} \Rightarrow m \cdot g \cdot h = (28.00 \times 10^{-3} \text{ kg})(9.8 \text{ m/s}^2) \cdot h = 3742 \text{ J}$$

Solving for h , the altitude reached by an N_2 molecule is 13637 m or 13.6 km.