

Assignment 6

Chapter 5: #82, 83, 84, 94, 114,88, 129 Ch. 12: #21,22

Kinetic Molecular Theory and Real Gases

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

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

83. a. $PV = nRT$

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

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

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

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

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

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

84. a. $PV = nRT$

$$P = \frac{nRT}{V} = \frac{0.5000 \text{ mol} \times \frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298.2 \text{ K}}{10.000 \text{ L}} = 1.224 \text{ atm}$$

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

$$\left[P + 1.39\left(\frac{0.5000}{10.000}\right)^2 \text{ atm} \right] (10.000 \text{ L} - 0.5000 \times 0.0391 \text{ L}) = 12.24 \text{ L atm}$$

$$(P + 0.00348 \text{ atm})(10.000 \text{ L} - 0.0196 \text{ L}) = 12.24 \text{ L atm}$$

$$P + 0.00348 \text{ atm} = \frac{12.24 \text{ L atm}}{9.980 \text{ L}} = 1.226 \text{ atm}, \quad P = 1.226 - 0.00348 = 1.223 \text{ atm}$$

- c. The results agree to $\pm 0.001 \text{ atm}$ (0.08%).
- d. In Exercise 83 the pressure is relatively high and there is significant disagreement. In Exercise 84 the pressure is around 1 atm and both gas laws show better agreement. The ideal gas law is valid at relatively low pressures.

94. $\frac{\text{Diffusion rate } ^{235}\text{UF}_6}{\text{Diffusion rate } ^{238}\text{UF}_6} = 1.0043$ (See Section 5.7 of the text.)

$$\frac{^{235}\text{UF}_6}{^{238}\text{UF}_6} \times (1.0043)^{100} = \frac{1526}{1.000 \times 10^5 - 1526}, \quad \frac{^{235}\text{UF}_6}{^{238}\text{UF}_6} \times 1.5358 = \frac{1526}{98500}$$

$$\frac{^{235}\text{UF}_6}{^{238}\text{UF}_6} = 1.01 \times 10^{-2} = \text{initial } ^{235}\text{U to } ^{238}\text{U atom ratio}$$

Additional Exercises

114. $33.5 \text{ mg CO}_2 \times \frac{12.01 \text{ mg C}}{44.01 \text{ mg CO}_2} = 9.14 \text{ mg C}; \quad \% \text{ C} = \frac{9.14 \text{ mg}}{35.0 \text{ mg}} \times 100 = 26.1\% \text{ C}$

$$41.1 \text{ mg H}_2\text{O} \times \frac{2.016 \text{ mg H}}{18.02 \text{ mg H}_2\text{O}} = 4.60 \text{ mg H}; \quad \% \text{ H} = \frac{4.60 \text{ mg}}{35.0 \text{ mg}} \times 100 = 13.1\% \text{ H}$$

$$n_{\text{N}_2} = \frac{P_{\text{N}_2} V}{RT} = \frac{\frac{740}{760} \text{ atm} \times 35.6 \times 10^{-3} \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{K mol}} \times 298 \text{ K}} = 1.42 \times 10^{-3} \text{ mol N}_2$$

$$1.42 \times 10^{-3} \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{\text{mol N}_2} = 3.98 \times 10^{-2} \text{ g nitrogen} = 39.8 \text{ mg nitrogen}$$

$$\text{Mass \% N} = \frac{39.8 \text{ mg}}{65.2 \text{ mg}} \times 100 = 61.0\% \text{ N}$$

Or we can get % N by difference: $\% \text{ N} = 100.0 - (26.1 + 13.1) = 60.8\%$

Out of 100.0 g:

$$26.1 \text{ g C} \times \frac{1 \text{ mol}}{12.01 \text{ g}} = 2.17 \text{ mol C}; \quad \frac{2.17}{2.17} = 1.00$$

$$13.1 \text{ g H} \times \frac{1 \text{ mol}}{1.008 \text{ g}} = 13.0 \text{ mol H}; \quad \frac{13.0}{2.17} = 5.99$$

$$60.8 \text{ g N} \times \frac{1 \text{ mol}}{14.01 \text{ g}} = 4.34 \text{ mol N}; \quad \frac{4.34}{2.17} = 2.00$$

Empirical formula is CH_6N_2 .

$$\frac{\text{Rate}_1}{\text{Rate}_2} = \left(\frac{M}{39.95} \right)^{1/2} = \frac{26.4}{24.6} = 1.07, \quad M = (1.07)^2 \times 39.95 = 45.7 \text{ g/mol}$$

Empirical formula mass of $\text{CH}_6\text{N}_2 \approx 12 + 6 + 28 = 46$. Thus the molecular formula is also CH_6N_2 .

88. The van der Waals constant b is a measure of the size of the molecule. Thus C_3H_8 should have the largest value of b because it has the largest molar mass (size).

$$21. \quad v = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{1.0 \times 10^{-2} \text{ m}} = 3.0 \times 10^{10} \text{ s}^{-1}$$

$$E = hv = 6.63 \times 10^{-34} \text{ J s} \times 3.0 \times 10^{10} \text{ s}^{-1} = 2.0 \times 10^{-23} \text{ J/photon}$$

$$\frac{2.0 \times 10^{-23} \text{ J}}{\text{photon}} = \frac{6.02 \times 10^{23} \text{ photons}}{\text{mol}} = 12 \text{ J/mol}$$

22. The wavelength is the distance between consecutive wave peaks. Wave a shows 4 wavelengths and wave b shows 8 wavelengths.

$$\text{Wave } a: \lambda = \frac{1.6 \times 10^{-3} \text{ m}}{4} = 4.0 \times 10^{-4} \text{ m}$$

$$\text{Wave } b: \lambda = \frac{1.6 \times 10^{-3} \text{ m}}{8} = 2.0 \times 10^{-4} \text{ m}$$

Wave a has the longer wavelength. Because frequency and photon energy are both inversely proportional to wavelength, wave b will have the higher frequency and larger photon energy since it has the shorter wavelength.

$$v = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{2.0 \times 10^{-4} \text{ m}} = 1.5 \times 10^{12} \text{ s}^{-1}$$

$$E = \frac{hc}{\lambda} = \frac{6.63 \times 10^{-34} \text{ J s} \times 3.00 \times 10^8 \text{ m/s}}{2.0 \times 10^{-4} \text{ m}} = 9.9 \times 10^{-22} \text{ J}$$

Because both waves are examples of electromagnetic radiation, both waves travel at the same speed, c , the speed of light. From Figure 12.3 of the text, both of these waves represent infrared electromagnetic radiation.

$$129. \quad \frac{PV}{nRT} = 1 + \beta P; \quad \frac{n}{V} \times \text{molar mass} = d$$

$$\frac{\text{molar mass}}{RT} \times \frac{P}{d} = 1 + \beta P, \quad \frac{P}{d} = \frac{RT}{\text{molar mass}} + \frac{\beta RTP}{\text{molar mass}}$$

This is in the equation for a straight line: $y = b + mx$. If we plot P/d versus P and extrapolate to $P = 0$, we get a y intercept = $b = 1.398 = RT/\text{molar mass}$.

$$\text{At } 0.00^\circ\text{C, molar mass} = \frac{0.08206 \times 273.15}{1.398} = 16.03 \text{ g/mol.}$$