

Assigned Problems**For Review**

7. a. At constant temperature, the average kinetic energy of the He gas sample will equal the average kinetic energy of the Cl<sub>2</sub> gas sample. In order for the average kinetic energies to be the same, the smaller He atoms must move at a faster average velocity as compared to Cl<sub>2</sub>. Therefore, plot A, with the slower average velocity, would be for the Cl<sub>2</sub> sample, and plot B would be for the He sample. Note the average velocity in each plot is a little past the top of the peak.
- b. As temperature increases, the average velocity of a gas will increase. Plot A would be for O<sub>2</sub>(g) at 273 K and plot B, with the faster average velocity, would be for O<sub>2</sub>(g) at 1273 K.

Because a gas behaves more ideally at higher temperatures, O<sub>2</sub>(g) at 1273 K would behave most ideally.

9. The pressure measured for real gases is too low as compared to ideal gases. This is due to the attractions gas particles do have for each other; these attractions “hold” them back from hitting the container walls as forcefully. To make up for this slight decrease in pressure for real gases, a factor is added to the measured pressure. The measured volume is too large. A fraction of the space of the container volume is taken up by the volume of gas of the molecules themselves. Therefore, the actual volume available to real gas molecules is slightly less than the container volume. A term is subtracted from the container volume to correct for the volume taken up by real gas molecules.

**Exercises****Pressure**

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

torr

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

$$29. \quad 6.5 \text{ cm} \times \frac{10 \text{ mm}}{\text{cm}} = 65 \text{ mm Hg} = 65 \text{ torr}; \quad 65 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 8.6 \times 10^{-2} \text{ atm}$$

$$8.6 \times 10^{-2} \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 8.7 \times 10^3 \text{ Pa}$$

31. If the levels of Hg in each arm of the manometer are equal, the pressure in the flask is equal to atmospheric pressure. When they are unequal, the difference in height in mm will be equal to the difference in pressure in mm Hg between the flask and the atmosphere. Which level is higher will tell us whether the pressure in the flask is less than or greater than atmospheric.

## CHAPTER 5 GASES

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a.  $P_{\text{flask}} < P_{\text{atm}}$ ;  $P_{\text{flask}} = 760. - 118 = 642 \text{ mm Hg} = 642 \text{ torr}$ ;  $642 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.845$   
atm

$$0.845 \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 8.56 \times 10^4 \text{ Pa}$$

b.  $P_{\text{flask}} > P_{\text{atm}}$ ;  $P_{\text{flask}} = 760. \text{ torr} + 215 \text{ torr} = 975 \text{ torr}$ ;  $975 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.28 \text{ atm}$

$$1.28 \text{ atm} \times \frac{1.013 \times 10^5 \text{ Pa}}{\text{atm}} = 1.30 \times 10^5 \text{ Pa}$$

c.  $P_{\text{flask}} = 635 - 118 = 517 \text{ torr}$ ;  $P_{\text{flask}} = 635 + 215 = 850. \text{ torr}$

### Gas Laws

35. At constant T and P, Avogadro's law holds.

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}, \quad n_2 = \frac{V_2 n_1}{V_1} = \frac{20. \text{ L} \times 0.50 \text{ mol}}{11.2 \text{ L}} = 0.89 \text{ mol}$$

As expected, as V increases, n increases.

39.  $n = \frac{PV}{RT} = \frac{135 \text{ atm} \times 200.0 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times (273 + 24) \text{ K}} = 1.11 \times 10^3 \text{ mol}$

For He:  $1.11 \times 10^3 \text{ mol} \times \frac{4.003 \text{ g He}}{\text{mol}} = 4.44 \times 10^3 \text{ g He}$

For H<sub>2</sub>:  $1.11 \times 10^3 \text{ mol} \times \frac{2.016 \text{ g He}}{\text{mol}} = 2.24 \times 10^3 \text{ g H}_2$

43. For a gas at two conditions:  $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$

Because V is constant:  $\frac{P_1}{n_1 T_1} = \frac{P_2}{n_2 T_2}$ ,  $n_2 = \frac{n_1 P_2 T_1}{P_1 T_2}$

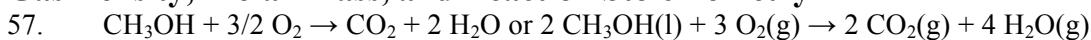
$$n_2 = \frac{1.50 \text{ mol} \times 800. \text{ torr} \times 298 \text{ K}}{400. \text{ torr} \times 323 \text{ K}} = 2.77 \text{ mol}$$

mol of gas added =  $n_2 - n_1 = 2.77 - 1.50 = 1.27 \text{ mol}$

For two condition problems, units for P and V just need to be the same units for both conditions, not necessarily atm and L. The unit conversions from other P or V units would cancel when applied to both conditions. However, temperature always must be converted

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to the Kelvin scale. The temperature conversions between other units and Kelvin will not cancel each other.

**Gas Density, Molar Mass, and Reaction Stoichiometry**

$$50.0 \text{ mL} \times \frac{0.850 \text{ g}}{\text{mL}} \times \frac{1 \text{ mol}}{32.04 \text{ g}} = 1.33 \text{ mol CH}_3\text{OH}(\text{l}) \text{ available}$$

$$n_{\text{O}_2} = \frac{PV}{RT} = \frac{2.00 \text{ atm} \times 22.8 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times 300. \text{ K}} = 1.85 \text{ mol O}_2 \text{ available}$$

$$1.33 \text{ mol CH}_3\text{OH} \times \frac{3 \text{ mol O}_2}{2 \text{ mol CH}_3\text{OH}} = 2.00 \text{ mol O}_2$$

2.00 mol O<sub>2</sub> are required to react completely with all of the CH<sub>3</sub>OH available. We only have 1.85 mol O<sub>2</sub>, so O<sub>2</sub> is limiting.

$$1.85 \text{ mol O}_2 \times \frac{4 \text{ mol H}_2\text{O}}{3 \text{ mol O}_2} = 2.47 \text{ mol H}_2\text{O}$$

61. molar mass =  $\frac{dRT}{P}$  where d = density of gas in units of g/L

$$\text{molar mass} = \frac{3.164 \text{ g/L} \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times 273.2 \text{ K}}{1.000 \text{ atm}} = 70.98 \text{ g/mol}$$

The gas is diatomic, so the atomic mass = 70.93/2 = 35.47. This is chlorine and the identity of the gas is Cl<sub>2</sub>.

63.  $d_{\text{UF}_6} = \frac{P \times (\text{molar mass})}{RT} = \frac{\left(745 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}\right) \times 352.0 \text{ g/mol}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times 333 \text{ K}} = 12.6 \text{ g/L}$

**Partial Pressure**

65.  $P_{\text{CO}_2} = \frac{nRT}{V} = \frac{\left(7.8 \text{ torr} \times \frac{1 \text{ mol}}{44.01 \text{ g}}\right) \times \frac{0.08206 \text{ L atm}}{\text{mol K}} \times 300. \text{ K}}{4.0 \text{ L}} = 1.1 \text{ atm}$

With air present, the partial pressure of CO<sub>2</sub> will still be 1.1 atm. The total pressure will be the sum of the partial pressures,  $P_{\text{total}} = P_{\text{CO}_2} + P_{\text{air}}$ .

$$P_{\text{total}} = 1.1 \text{ atm} + \left(740 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}\right) = 1.1 + 0.97 = 2.1 \text{ atm}$$

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

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

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

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

69. a. mol fraction  $\text{CH}_4 = \chi_{\text{CH}_4} = \frac{P_{\text{CH}_4}}{P_{\text{total}}} = \frac{0.175 \text{ atm}}{0.175 \text{ atm} + 0.250 \text{ atm}} = 0.412$

$$\chi_{\text{O}_2} = 1.000 - 0.412 = 0.588$$

b.  $PV = nRT$ ,  $n_{\text{total}} = \frac{P_{\text{total}} \times V}{RT} = \frac{0.425 \text{ atm} \times 10.5 \text{ L}}{\frac{0.08206 \text{ L atm}}{\text{mol K}} \times 338 \text{ K}} = 0.161 \text{ mol}$

c.  $\chi_{\text{CH}_4} = \frac{n_{\text{CH}_4}}{n_{\text{total}}}$ ,  $n_{\text{CH}_4} = \chi_{\text{CH}_4} \times n_{\text{total}} = 0.412 \times 0.161 \text{ mol} = 6.63 \times 10^{-2} \text{ mol CH}_4$

$$6.63 \times 10^{-2} \text{ mol CH}_4 \times \frac{16.04 \text{ g CH}_4}{\text{mol CH}_4} = 1.06 \text{ g CH}_4$$

$$n_{\text{O}_2} = 0.588 \times 0.161 \text{ mol} = 9.47 \times 10^{-2} \text{ mol O}_2; 9.47 \times 10^{-2} \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{\text{mol O}_2} = 3.03 \text{ g O}_2$$

71.  $P_{\text{TOT}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}}$ ,  $1.032 \text{ atm} = \text{Error! Objects cannot be created from editing field codes.} +$

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

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

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

75.  $2 \text{ HN}_3(\text{g}) \rightarrow 3 \text{ N}_2(\text{g}) + \text{H}_2(\text{g})$ ; At constant V and T, P is directly proportional to n. In the reaction, we go from 2 moles of gaseous reactants to 4 moles of gaseous products. Since moles doubled, the final pressure will double ( $P_{\text{tot}} = 6.0 \text{ atm}$ ). Similarly, from the 2:1 mole ratio between  $\text{HN}_3$  and  $\text{H}_2$ , the partial pressure of  $\text{H}_2$  will be  $3.0/2 = 1.5 \text{ atm}$ . The partial pressure of  $\text{N}_2$  will be  $3/2$  ( $3.0 \text{ atm}$ ) =  $4.5 \text{ atm}$ . This is from the 2:3 mole ratio between  $\text{HN}_3$  and  $\text{N}_2$ .

Assigned Problems**Kinetic Molecular Theory and Real Gases**

77.  $KE_{\text{avg}} = 3/2 RT$ ; The average kinetic energy depends only on temperature. At each temperature,  $\text{CH}_4$  and  $\text{N}_2$  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,  $6.022 \times 10^{23}$  molecules/mol.

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

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

79.  $u_{\text{rms}} = \left( \frac{3RT}{M} \right)^{1/2}$ ,  $R = \frac{8.3145 \text{ J}}{\text{mol K}}$  and  $M = \text{molar mass in kg} = 1.604 \times 10^{-2} \text{ kg/mol}$  for  $\text{CH}_4$

$$\text{For } \text{CH}_4 \text{ at } 273 \text{ K: } u_{\text{rms}} = \left( \frac{\frac{3 \times 8.3145 \text{ J}}{\text{mol K}} \times 273 \text{ K}}{1.604 \times 10^{-2} \text{ kg/mol}} \right)^{1/2} = 652 \text{ m/s}$$

Similarly,  $u_{\text{rms}}$  for  $\text{CH}_4$  at 546 K is 921 m/s.

$$\text{For } \text{N}_2 \text{ at } 273 \text{ K: } u_{\text{rms}} = \left( \frac{\frac{3 \times 8.3145 \text{ J}}{\text{mol K}} \times 273 \text{ K}}{2.802 \times 10^{-2} \text{ kg/mol}} \right)^{1/2} = 493 \text{ m/s}$$

Similarly, for  $\text{N}_2$  at 546 K,  $u_{\text{rms}} = 697 \text{ m/s}$ .

81.	a	b	c	d
avg. KE	inc	dec	same (KE % T)	same
avg. velocity	inc	dec	same ( $\frac{1}{2} mv^2 = \text{KE} \% T$ )	same
coll. freq wall	inc	dec	inc	inc

Average kinetic energy and average velocity depend on T. As T increases, both average kinetic energy and average velocity increase. At constant T, both average kinetic energy and average velocity are constant. The collision frequency is proportional to the average velocity (as velocity increases it takes less time to move to the next collision) and to the quantity  $n/V$  (as molecules per volume increase, collision frequency increases).

83. a. They will all have the same average kinetic energy since they are all at the same temperature.
- b. Flask C;  $\text{H}_2$  has the smallest molar mass. At constant T, the lightest molecules are the fastest (on the average). This must be true in order for the average kinetic energies to be constant.