

REVIEW QUESTIONS

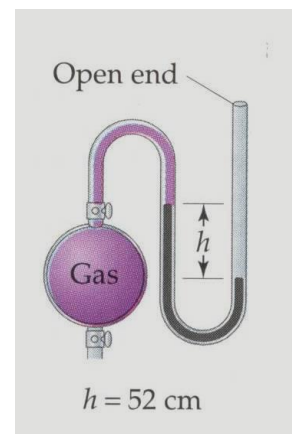
## Chapter 5

1. Determine the pressure of the gas (in mmHg) in the diagram below, given atmospheric pressure = 0.975 atm.

$$P_{\text{atm}} = 0.975 \text{ atm} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 741 \text{ mmHg}$$

$$h = 52 \text{ cmHg} \times \frac{10 \text{ mmHg}}{1 \text{ cmHg}} = 520 \text{ mmHg}$$

$$P_{\text{gas}} = P_{\text{atm}} - h = 741 \text{ mmHg} - 520 \text{ mmHg} = 221 \text{ mmHg}$$



2. A sample of oxygen gas has a volume of 26.7 L at 752 mmHg and 20°C. What is the volume of this gas at 1.30 atm and 20°C?

$$V_2 = V_1 \times \frac{P_1}{P_2} = 26.7 \text{ L} \times \frac{752 \text{ mmHg}}{1.30 \text{ atm} \times \frac{760 \text{ mmHg}}{1 \text{ atm}}} = 20.3 \text{ L}$$

3. A 35.8 L cylinder of Argon gas is connected to and transferred into an evacuated 1875-L tank at constant temperature. If the final pressure in the tank is 721 mmHg, what must have been the original pressure (in atm) in the cylinder?

$$P_1 = P_2 \times \frac{V_2}{V_1} = (721 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}) \times \frac{1875 \text{ L}}{35.8 \text{ L}} = 49.7 \text{ atm}$$

4. A 34.0-L cylinder contains 305 g of oxygen gas at 22°C. How many grams of gas must be released to reduce the pressure in the cylinder to 1.15 atm if the temperature remains constant?

Mass of gas present at 1.15 atm

$$n = \frac{PV}{RT} = \frac{(1.15 \text{ atm})(34.0 \text{ L})}{(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(295 \text{ K})} = 1.614 \text{ mol}$$

$$\text{mass of oxygen} = 1.614 \text{ mol} \times \frac{32.0 \text{ g}}{1 \text{ mol}} = 51.6 \text{ g}$$

Mass of gas that must be removed

$$305 \text{ g} - 51.6 \text{ g} = 253 \text{ g}$$

5. At STP, 0.280 L of a gas weighs 0.400 g. Calculate the molar mass of this gas.

$$\text{moles of gas} = 0.280 \cancel{\text{ L}} \times \frac{1 \text{ mol}}{22.4 \cancel{\text{ L}}} = 0.0125 \text{ mol}$$

$$\text{molar mass} = \frac{0.400 \text{ g}}{0.0125 \text{ mol}} = 32.0 \text{ g/mol}$$

6. Calculate the density of HBr gas in g/L at 733 mmHg and 46°C.

$$d = \frac{PM}{RT} = \frac{\left(\frac{733}{760} \text{ atm}\right)(80.91 \text{ g/mol})}{(0.0821 \text{ Latm/molK})(319 \text{ K})} = 2.98 \text{ g/L}$$

7. A mixture of 4.00 g of hydrogen and 10.0 g of helium are in a 4.30-L flask at 0°C. What is the total pressure of the container and the partial pressures of each gas?

$$\text{mol H}_2 = 4.00 \cancel{\text{ g}} \times \frac{1 \text{ mol}}{2.02 \cancel{\text{ g}}} = 1.98 \text{ mol}$$

$$\text{mol He} = 10.0 \cancel{\text{ g}} \times \frac{1 \text{ mol}}{4.00 \cancel{\text{ g}}} = 2.50 \text{ mol}$$

$$\text{Total mol} = 1.98 + 2.50 = 4.48 \text{ mol}$$

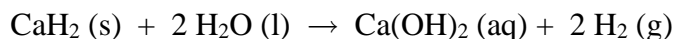
$$P_{\text{Total}} = \frac{nRT}{V} = \frac{(4.48 \cancel{\text{ mol}})(0.0821 \frac{\cancel{\text{ L}} \text{ atm}}{\cancel{\text{ mol}} \text{ K}})(273 \cancel{\text{ K}})}{4.30 \cancel{\text{ L}}} = 23.4 \text{ atm}$$

$$X_{\text{H}_2} = \frac{1.98 \text{ mol}}{4.48 \text{ mol}} = 0.442 \quad X_{\text{He}} = \frac{2.50 \text{ mol}}{4.48 \text{ mol}} = 0.558$$

$$P_{\text{H}_2} = X_{\text{H}_2} P_{\text{Tot}} = (0.442)(23.4 \text{ atm}) = 10.3 \text{ atm}$$

$$P_{\text{He}} = X_{\text{He}} P_{\text{Tot}} = (0.558)(23.4 \text{ atm}) = 13.1 \text{ atm}$$

8. Life rafts and weather balloons can be inflated by the reaction shown below:

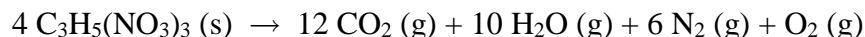


How many grams of  $\text{CaH}_2$  are needed to produce 10.0 L of hydrogen gas at 740 mmHg and 23°C?

$$n_{\text{H}_2} = \frac{PV}{RT} = \frac{(740 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}})(10.0 \text{ L})}{(0.0821 \frac{\text{L atm}}{\text{mol K}})(296 \text{ K})} = 0.4007 \text{ mol}$$

$$0.4007 \text{ mol H}_2 \times \frac{1 \text{ mol CaH}_2}{2 \text{ mol H}_2} \times \frac{42.10 \text{ g}}{1 \text{ mol}} = 8.43 \text{ g CaH}_2$$

9. Nitroglycerin, an explosive compound, decomposes according to the equation below:



Calculate the total volume of gases produced at 1.2 atm and 26°C when 260 g of nitroglycerine is decomposed.

$$260 \text{ g NG} \times \frac{1 \text{ mol}}{227.1 \text{ g}} \times \frac{29 \text{ mol gas}}{4 \text{ mol NG}} = 8.30 \text{ mol gas}$$

$$V = \frac{nRT}{P} = \frac{(8.30 \text{ mol})(0.0821 \frac{\text{L atm}}{\text{mol K}})(299 \text{ K})}{1.2 \text{ atm}} = 170 \text{ L}$$

10. A 1.65-g sample of Al is reacted with excess HCl and the hydrogen produced is collected over water at 25°C at a barometric pressure of 744 mmHg. What volume of hydrogen gas is produced in this reaction? (Vapor pressure of water at 25°C is 23.8 mmHg)

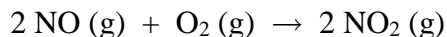


$$P_{\text{H}_2} = P_{\text{atm}} - P_{\text{H}_2\text{O}} = 744 \text{ mmHg} - 23.8 \text{ mmHg} = 720.2 \text{ mmHg}$$

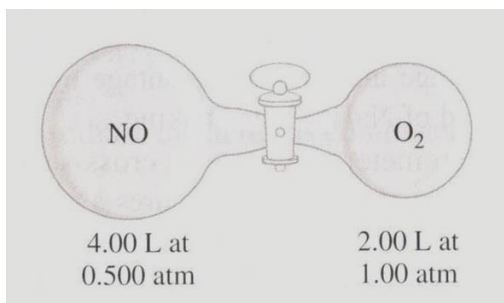
$$1.65 \text{ g Al} \times \frac{1 \text{ mol}}{27.0 \text{ g}} \times \frac{3 \text{ mol H}_2}{2 \text{ mol Al}} = 0.09173 \text{ mol H}_2$$

$$V = \frac{nRT}{P} = \frac{(0.09173 \text{ mol})(0.0821 \frac{\text{L atm}}{\text{mol K}})(298 \text{ K})}{720.2 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}} = 2.37 \text{ L}$$

11. Nitric oxide reacts with oxygen gas as shown below:



Initially NO and O<sub>2</sub> are separated as shown in the diagram below. When the valve is opened the reaction quickly goes to completion. Determine the identity of the gases that remain at the end of the reaction and their partial pressure. Assume temperature remains at 25°C.



Since the system is at constant temperature, the moles of gas is proportional to the product of pressure and volume as shown below:

$$PV = nRT$$

at constant temp.

$$PV \propto n$$

Before reaction

$$\text{mol NO} \propto (4.00 \text{ L})(0.500 \text{ atm}) = 2.00$$

$$\text{mol O}_2 \propto (2.00 \text{ L})(1.00 \text{ atm}) = 2.00$$

	$2 \text{NO}$	$+ \text{O}_2$	$\longrightarrow$	$2 \text{NO}_2$
<b>Initial</b>	<b>2.00</b>	<b>2.00</b>		<b>0</b>
<b>React</b>	<b>2.00</b>	<b>1.00</b>		<b>2.00</b>
<b>End</b>	<b>0</b>	<b>1.00</b>		<b>2.00</b>

At end of reaction, total volume is 6.00 L. Therefore,

$$P_{\text{O}_2} \propto \frac{\text{mol}}{V} = \frac{1.00}{6.00} = 0.167 \text{ atm}$$

$$P_{\text{NO}_2} \propto \frac{\text{mol}}{V} = \frac{2.00}{6.00} = 0.333 \text{ atm}$$

12. A 4.85-g sample of solid ammonium chloride is placed in a 1.50-L evacuated flask and heated until it decomposes, as shown below:



After the reaction has completed, the total pressure in the flask is measured as 4.40 atm at 202°C. Based on this information, what percent of the ammonium chloride decomposed?

$$P_{\text{total}} = P_{\text{NH}_3} + P_{\text{HCl}} \quad \text{and} \quad P_{\text{NH}_3} = P_{\text{HCl}}$$

$$P_{\text{NH}_3} = \frac{4.40 \text{ atm}}{2} = 2.20 \text{ atm}$$

$$n_{\text{NH}_3} = \frac{PV}{RT} = \frac{(2.20 \text{ atm})(1.50 \text{ L})}{(0.0821 \frac{\text{L atm}}{\text{mol K}})(475 \text{ K})} = 0.0846 \text{ mol}$$

$$0.0846 \text{ mol NH}_3 \times \frac{1 \text{ mol NH}_4\text{Cl}}{1 \text{ mol NH}_3} \times \frac{53.50 \text{ g}}{1 \text{ mol}} = 4.526 \text{ g NH}_4\text{Cl reacted}$$

$$\% \text{ decomposed} = \frac{4.526 \text{ g}}{4.85 \text{ g}} \times 100 = 93.3\%$$

13. A mixture of gases containing 12.45 g of H<sub>2</sub>, 60.67 g of N<sub>2</sub> and 2.380 g of NH<sub>3</sub> are placed in a 10.00 L container at 90°C. What is the total pressure (in atm) and partial pressure of each component in the gas mixture?

$$\text{mol H}_2 = 12.45 \text{ g} \times \frac{1 \text{ mol}}{2.016 \text{ g}} = 6.176 \text{ mol}$$

$$\text{mol N}_2 = 60.67 \text{ g} \times \frac{1 \text{ mol}}{28.02 \text{ g}} = 2.165 \text{ mol}$$

$$\text{mol NH}_3 = 2.380 \text{ g} \times \frac{1 \text{ mol}}{17.04 \text{ g}} = 0.1397 \text{ mol}$$

$$\text{Total mol} = 6.176 + 2.165 + 0.1397 = 8.481 \text{ mol}$$

$$P_{\text{Total}} = \frac{nRT}{V} = \frac{(8.481 \text{ mol})(0.0821 \frac{\text{L atm}}{\text{mol K}})(363 \text{ K})}{10.00 \text{ L}} = 25.28 \text{ atm}$$

$$X_{\text{H}_2} = \frac{6.176 \text{ mol}}{8.481 \text{ mol}} = 0.7282 \quad P_{\text{H}_2} = X_{\text{H}_2} P_{\text{Tot}} = (0.7282)(25.28 \text{ atm}) = 18.41 \text{ atm}$$

$$X_{\text{N}_2} = \frac{2.165 \text{ mol}}{8.481 \text{ mol}} = 0.2553 \quad P_{\text{N}_2} = X_{\text{N}_2} P_{\text{Tot}} = (0.2553)(25.28 \text{ atm}) = 6.454 \text{ atm}$$

$$X_{\text{NH}_3} = \frac{0.1397 \text{ mol}}{8.481 \text{ mol}} = 0.01647 \quad P_{\text{NH}_3} = X_{\text{NH}_3} P_{\text{Tot}} = (0.01647)(25.28 \text{ atm}) = 0.4164 \text{ atm}$$

14. Which has a higher average speed, H<sub>2</sub> at 150 K or He at 375°C?

$$u_{\text{H}_2} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 (8.314 \text{ J/molK})(150 \text{ K})}{2.02 \times 10^{-3} \text{ kg/mol}}} = 1360 \text{ m/s}$$

$$u_{\text{He}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 (8.314 \text{ J/molK})(648 \text{ K})}{4.00 \times 10^{-3} \text{ kg/mol}}} = 2010 \text{ m/s}$$

**He at 375 °C has higher average speed**

15. A big-league fastball travels at about 45.0 m/s. At what temperature (°C) do helium atoms have this same average speed?

$$u_{\text{He}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 (8.314 \text{ J/molK})(T)}{4.00 \times 10^{-3} \text{ kg/mol}}} = 45.0 \text{ m/s}$$

**squaring both sides,**

$$\frac{3 (8.314 \text{ J/molK})(T)}{4.00 \times 10^{-3} \text{ kg/mol}} = 2025 \text{ m}^2/\text{s}^2$$

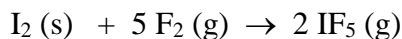
$$T = 0.325 \text{ K}$$

$$T = 0.325 - 273.15 = -272.83 \text{ °C}$$

16. The surface temperature of Venus is about 1050 K, and the pressure is about 75.0 earth atmospheres. Assuming that these conditions represent a Venusian “STP”, what is the “standard” molar volume of a gas on Venus?

$$V = \frac{nRT}{P} = \frac{(1.00 \text{ mol})(0.0821 \frac{\text{L atm}}{\text{mol K}})(1050 \text{ K})}{75.0 \text{ atm}} = 1.15 \text{ L}$$

17. Gaseous iodine IF<sub>5</sub> can be prepared by the reaction of solid iodine and gaseous fluorine:



A 5.00-L flask is charged with 10.0 g of I<sub>2</sub> and 10.0 g of F<sub>2</sub>, and the reaction proceeds until one of the reagents is completely consumed. After the reaction is complete the temperature in the flask is 125°C. What is the partial pressure of IF<sub>5</sub> in the flask?

**Before reaction,**

$$\text{mol I}_2 = 10.0 \text{ g} \times \frac{1 \text{ mol}}{253.8 \text{ g}} = 0.0394 \text{ mol}$$

$$\text{mol F}_2 = 10.0 \text{ g} \times \frac{1 \text{ mol}}{38.00 \text{ g}} = 0.263 \text{ mol}$$

	$\text{I}_2$	$+ 5 \text{F}_2$	$\longrightarrow$	$2 \text{IF}_5$
<b>Initial</b>	<b>0.0394</b>	<b>0.263</b>		<b>0</b>
<b>React</b>	<b>0.0394</b>	<b>0.197</b>		<b>0.0788</b>
<b>End</b>	<b>0</b>	<b>0.066</b>		<b>0.0788</b>

**At end of reaction,**

$$\text{Total moles} = 0.066 + 0.0788 = 0.1448 \text{ mol}$$

$$P_{\text{total}} = \frac{n_{\text{total}}RT}{V} = \frac{(0.1448 \text{ mol})(0.821 \text{ Latm/molK})(398 \text{ K})}{5.00 \text{ L}} = 0.946 \text{ atm}$$

$$X_{\text{IF}_5} = \frac{0.0788 \text{ mol}}{0.1448 \text{ mol}} = 0.544 \quad P_{\text{IF}_5} = X_{\text{IF}_5} P_{\text{total}} = (0.544)(0.946 \text{ atm}) = 0.515 \text{ atm}$$

18. A sample of N<sub>2</sub>O effuses from a container in 42.0 seconds. How long will it take the same amount of gaseous I<sub>2</sub> to effuse from the same container under identical conditions?

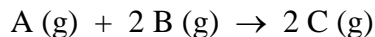
**Rate of effusion can be calculated as inverse of effusion time. Therefore,**

$$\frac{\text{Rate N}_2\text{O}}{\text{Rate I}_2} = \sqrt{\frac{M_{\text{I}_2}}{M_{\text{N}_2\text{O}}}} = \sqrt{\frac{254 \text{ g/mol}}{44.02}} = 2.40$$

$$\frac{\text{Rate N}_2\text{O}}{\text{Rate I}_2} = \frac{1/T_{\text{N}_2\text{O}}}{1/T_{\text{I}_2}} = \frac{T_{\text{I}_2}}{T_{\text{N}_2\text{O}}} = 2.40$$

$$T_{\text{I}_2} = (T_{\text{N}_2\text{O}}) (2.40) = (42.0 \text{ s})(2.40) = 101 \text{ s}$$

19. This reaction occurs in a closed container:



A reaction mixture initially contains 1.5 L of A and 2.0 L of B. Assuming that the volume and temperature of the reaction mixture is constant, how does the pressure change, and what is the percentage change if the reaction goes to completion.

**Since temperature is constant, the volumes of each reactant is proportional to the number of moles present for each. Therefore, the following reaction table can be set up to track the progress of the reaction from beginning to end:**

	<b>A (g)</b>	<b>+</b>	<b>2 B (g)</b>	<b>→</b>	<b>2 C (g)</b>
<b>Initial</b>	<b>1.5 mol</b>		<b>2.0 mol</b>		<b>0</b>
<b>React</b>	<b>1.0 mol</b>		<b>2.0 mol</b>		<b>2.0 mol</b>
<b>End</b>	<b>0.5 mol</b>		<b>0</b>		<b>2.0 mol</b>

**Since the volume of flask is constant, moles and pressure are proportional. Therefore, after reaction, pressure decreases since there are less moles of gas present than at the start.**

$$\% \text{ change in pressure} = \frac{2.5 - 3.5}{3.5} \times 100 = -29\%$$

20. A gas mixture contains 30.0% CH<sub>4</sub> and 70.0% Xe by mass. If the total pressure of this mixture is 0.44 atm, what is the partial pressure of each gas in the mixture?

**Assume 100 g of gas mixture:**

$$\text{mol CH}_4 = 30.0 \text{ g} \times \frac{1 \text{ mol}}{16.05 \text{ g}} = 1.87 \text{ mol}$$

$$\text{mol Xe} = 70.0 \text{ g} \times \frac{1 \text{ mol}}{131.29 \text{ g}} = 0.533 \text{ mol}$$

$$\text{Total mol} = 1.87 + 0.533 = 2.403 \text{ mol}$$

$$X_{\text{CH}_4} = \frac{1.87 \text{ mol}}{2.403 \text{ mol}} = 0.778 \quad P_{\text{CH}_4} = X_{\text{CH}_4} P_{\text{Tot}} = (0.778)(0.44 \text{ atm}) = 0.34 \text{ atm}$$

$$X_{\text{Xe}} = \frac{0.533 \text{ mol}}{2.403 \text{ mol}} = 0.222 \quad P_{\text{Xe}} = X_{\text{Xe}} P_{\text{Tot}} = (0.222)(0.44 \text{ atm}) = 0.10 \text{ atm}$$