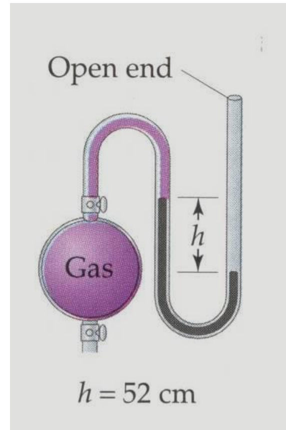
Chemistry 101 **ANSWER KEY**

REVIEW QUESTIONS
Chapter 5

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



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

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

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

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?

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

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?

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

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

$$n = \frac{PV}{RT} = \frac{(1.15\ \cancel{atm})(34.0\ \cancel{L})}{\left(0.0821 \frac{\cancel{L\ atm}}{mol\ \cancel{K}}\right)(295\ \cancel{K})} = 1.6\underline{1}4\ mol$$

$$\text{mass of oxygen} = 1.6\underline{1}4\ \cancel{mol} \times \frac{32.0\ g}{1\ \cancel{mol}} = 51.6\ g$$

Mass of gas that must be removed

$$305\ g - 51.6\ g = 253\ g$$

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

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

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

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

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

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?

$$\text{mol } H_2 = 4.00\ \cancel{g} \times \frac{1\ \text{mol}}{2.02\ \cancel{g}} = 1.98\ \text{mol}$$

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

Total mol = 1.98 + 2.50 = 4.48 mol

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

$$X_{H_2} = \frac{1.98\ \text{mol}}{4.48\ \text{mol}} = 0.442 \qquad X_{He} = \frac{2.50\ \text{mol}}{4.48\ \text{mol}} = 0.558$$

$$P_{H_2} = X_{H_2}P_{Tot} = (0.442)(23.4\ \text{atm}) = 10.3\ \text{atm}$$

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

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

$$CaH_2\,(s) + 2\,H_2O\,(l) \rightarrow Ca(OH)_2\,(aq) + 2\,H_2\,(g)$$

How many grams of CaH_2 are needed to produce 10.0 L of hydrogen gas at 740 mmHg and 23°C?

$$n_{H_2} = \frac{PV}{RT} = \frac{(740\ \cancel{mmHg}\ x\ \frac{1\ \cancel{atm}}{760\ \cancel{mmHg}})(10.0\ \cancel{L})}{(0.0821\frac{\cancel{L\ atm}}{mol\ \cancel{K}})(296\ \cancel{K})} = 0.40\underline{0}7\ mol$$

$$0.40\underline{0}7\ \cancel{mol\ H_2}\ x\ \frac{1\ \cancel{mol}\ CaH_2}{2\ \cancel{mol\ H_2}}\ x\ \frac{42.10\ g}{1\ \cancel{mol}} = 8.43\ g\ CaH_2$$

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

$$4\,C_3H_5(NO_3)_3\,(s) \rightarrow 12\,CO_2\,(g) + 10\,H_2O\,(g) + 6\,N_2\,(g) + O_2\,(g)$$

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

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

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

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)

$$2\,Al\,(s) + 6\,HCl\,(aq) \rightarrow 2\,AlCl_3\,(aq) + 3\,H_2\,(g)$$

$$P_{H_2} = P_{atm} - P_{H_2O} = 744\ mmHg - 23.8\ mmHg = 72\underline{0}.2\ mmHg$$

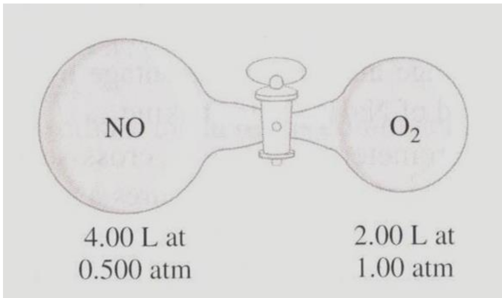
$$1.65\ \cancel{g\ Al}\ x\ \frac{1\ \cancel{mol}}{27.0\ \cancel{g}}\ x\ \frac{3\ mol\ H_2}{2\ \cancel{mol\ Al}} = 0.091\underline{7}3\ mol\ H_2$$

$$V = \frac{nRT}{P} = \frac{(0.091\underline{7}3\ \cancel{mol})(0.0821\frac{L\ \cancel{atm}}{\cancel{mol\ K}})(298\ \cancel{K})}{72\underline{0}.2\ \cancel{mmHg}\ x\ \frac{1\ atm}{760\ \cancel{mmHg}}} = 2.37\ L$$

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

$$2\,NO\,(g) + O_2\,(g) \rightarrow 2\,NO_2\,(g)$$

Initially NO and O_2 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:

$\mathbf{PV = n\,\cancel{RT}}$ **at constant temp.** $\mathbf{PV \propto n}$

Before reaction

mol NO α (4.00 L)(0.500 atm) = 2.00

mol O_2 α (2.00 L)(1.00 atm) = 2.00

$$2\,NO + O_2 \longrightarrow 2\,NO_2$$

Since 2 moles of each reactant are present at the start of reaction, NO is the limiting reactant and after the reaction is complete 1 mol of O_2 and 2 mol of NO_2 are present. At end of reaction, total volume is 6.00 L. Therefore,

$$P_{O_2}\ \alpha\ \frac{mol}{V} = \frac{1.00}{6.00} = 0.167\ atm$$

$$P_{NO_2}\ \alpha\ \frac{mol}{V} = \frac{2.00}{6.00} = 0.333\ atm$$

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:

$$NH_4Cl\,(s) \rightarrow NH_3\,(g) + HCl\,(g)$$

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?

$$P_{total} = P_{NH3} + P_{HCl} \quad \text{and} \quad P_{NH3} = P_{HCl}$$

$$P_{NH3} = \frac{4.40\text{ atm}}{2} = 2.20\text{ atm}$$

$$n_{NH3} = \frac{PV}{RT} = \frac{(2.20\ \cancel{\text{atm}})(1.50\ \cancel{\text{L}})}{(0.0821\frac{\cancel{\text{L}}\ \text{atm}}{\cancel{\text{mol}}\ \cancel{\text{K}}})(475\ \cancel{\text{K}})} = 0.0846\text{ mol}$$

$$0.0846\ \cancel{\text{mol NH}_3} \times \frac{1\text{ mol }NH_4Cl}{1\ \cancel{\text{mol NH}_3}} \times \frac{53.50\text{ g}}{1\text{ mol}} = 4.5\underline{2}6\text{ g }NH_4Cl\text{ reacted}$$

$$\%\text{ decomposed} = \frac{4.5\underline{2}6\text{ g}}{4.85\text{ g}} \times 100 = 93.3\%$$

13. A mixture of gases containing 12.45 g of H_2, 60.67 g of N_2 and 2.380 g of NH_3 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?

$$\text{mol }H_2 = 12.45\ \cancel{\text{g}} \times \frac{1\text{ mol}}{2.016\ \cancel{\text{g}}} = 6.176\text{ mol}$$

$$\text{mol }N_2 = 60.67\ \cancel{\text{g}} \times \frac{1\text{ mol}}{28.02\ \cancel{\text{g}}} = 2.165\text{ mol}$$

$$\text{mol }NH_3 = 2.380\ \cancel{\text{g}} \times \frac{1\text{ mol}}{17.04\ \cancel{\text{g}}} = 0.1397\text{ mol}$$

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

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

$$X_{H_2} = \frac{6.176\text{ mol}}{8.481\text{ mol}} = 0.7282 \qquad P_{H_2} = X_{H_2}P_{Tot} = (0.7282)(25.28\text{ atm}) = 18.41\text{ atm}$$

$$X_{He} = \frac{2.165\text{ mol}}{8.481\text{ mol}} = 0.2553 \qquad P_{He} = X_{He}P_{Tot} = (0.2553)(25.28\text{ atm}) = 6.454\text{ atm}$$

$$X_{NH3} = \frac{0.1397\text{ mol}}{8.481\text{ mol}} = 0.01647 \qquad P_{NH3} = X_{NH3}P_{Tot} = (0.01647)(25.28\text{ atm}) = 0.4164\text{ atm}$$

14. Which has a higher average speed, H_2 at 150 K or He at 375°C?

$$u_{H2} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3\ (8.314\ J/molK)(150\ K)}{2.02x10^{-3}\ kg/mol}} = 1360\ m/s$$

$$u_{He} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3\ (8.314\ J/molK)(648\ K)}{4.00x10^{-3}\ kg/mol}} = 2010\ m/s$$

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?

$$u_{He} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3\ (8.314\ J/molK)(T)}{4.00x10^{-3}\ kg/mol}} = 45.0\ m/s$$

squaring both sides,

$$\frac{3\ (8.314\ J/molK)(T)}{4.00x10^{-3}\ kg/mol} = 2025\ m^2/s^2$$

T = 0.325 K

T = 0.325 -273.15 = -272.83 °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?

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

17. Gaseous iodine IF_5 can be prepared by the reaction of solid iodine and gaseous fluorine:

$$I_2\,(s) + 5\,F_2\,(g) \rightarrow 2\,IF_5\,(g)$$

A 5.00-L flask is charged with 10.0 g of I_2 and 10.0 g of F_2, 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_5 in the flask?

Before reaction,

$$\text{mol } I_2 = 10.0\text{ g x } \frac{1\text{ mol}}{253.8\text{ g}} = 0.0394\text{ mol}$$

$$\text{mol } F_2 = 10.0\text{ g x } \frac{1\text{ mol}}{38.00\text{ g}} = 0.263\text{ mol}$$

	I_2	+ $5\,F_2$	$\longrightarrow$ $2\,IF_5$
Initial	**0.0394**	**0.263**	**0**
React	**0.0394**	**0.197**	**0.0788**
End	**0**	**0.066**	**0.0788**

At end of reaction,

Total moles = 0.066 + 0.0788 = 0.14$\underline{4}$8 mol

$$P_{total} = \frac{n_{total}RT}{V} = \frac{(0.14\underline{4}8\text{ mol})(0.821\text{ Latm/molK})(398\text{ K})}{5.00\text{ L}} = 0.946\text{ atm}$$

$$X_{IF_5} = \frac{0.0788\text{ mol}}{0.14\underline{4}8\text{ mol}} = 0.544 \qquad P_{IF_5} = X_{IF_5}P_{total} = (0.544)(0.946\text{ atm}) = 0.515\text{ atm}$$

18. A sample of N_2O effuses from a container in 42.0 seconds. How long will it take the same amount of gaseous I_2 to effuse from the same container under identical conditions?

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

$$\frac{\text{Rate } N_2O}{\text{Rate } I_2} = \sqrt{\frac{M_{I_2}}{M_{N_2O}}} = \sqrt{\frac{254\text{ g/mol}}{44.02}} = 2.40$$

$$\frac{\text{Rate } N_2O}{\text{Rate } I_2} = \frac{1/T_{N_2O}}{1/T_{I_2}} = \frac{T_{I_2}}{T_{N_2O}} = 2.40$$

$$T_{I_2} = (T_{N_2O})\,(2.40) = (42.0\text{ s})(2.40) = 101\text{ s}$$

19. This reaction occurs in a closed container:

$$A\,(g) + 2\,B\,(g) \rightarrow 2\,C\,(g)$$

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:

	A (g) +	2 B (g)	$\longrightarrow$ 2 C (g)
Initial	**1.5 mol**	**2.0 mol**	**0**
React	**1.0 mol**	**2.0 mol**	**2.0 mol**
End	**0.5 mol**	**0**	**2.0 mol**

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_4 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:

$$\text{mol } CH_4 = 30.0\ \cancel{g} \times \frac{1\ \text{mol}}{16.05\ \cancel{g}} = 1.87\ \text{mol}$$

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

Total mol = 1.87 + 0.533 = 2.403 mol

$$X_{CH_4} = \frac{1.87\ \text{mol}}{2.403\ \text{mol}} = 0.778 \qquad P_{CH_4} = X_{CH_4} P_{Tot} = (0.778)(0.44\ \text{atm}) = 0.34\ \text{atm}$$

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