

## Gas Laws III

Name \_\_\_\_\_ Lab Section # \_\_\_\_\_

1. Calculate the volume of a sample of helium at  $-33.0\text{ }^{\circ}\text{C}$  and  $1.23\text{ atm}$  if it occupies a volume of  $2.34\text{ L}$  at  $54.5\text{ }^{\circ}\text{C}$  and  $1026\text{ mmHg}$ .

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{T_2}{P_2} \left( \frac{P_1V_1}{T_1} \right) = \frac{240\text{ K}}{1.23\text{ atm}} \left( \frac{1.35\text{ atm} \cdot 2.34\text{ L}}{327.5\text{ K}} \right) = 1.88\text{ L}$$

2. A  $0.751\text{ mol}$  sample of an ideal gas occupies a  $10.0\text{ liter}$  flask at  $27.0\text{ }^{\circ}\text{C}$  and  $1.85\text{ atm}$ . If  $0.257\text{ mol}$  of the gas are removed from the container, calculate the new pressure. (Assume the temperature remains constant.)

$$PV = nRT$$

$$P = \frac{nRT}{V}$$

$$P = \frac{0.494\text{ mol} (0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(300\text{ K})}{10.0\text{ L}} = 1.22\text{ atm}$$

3. What is the volume of a bulb that contains  $3.56\text{ g}$  of nitrogen gas at  $25.0\text{ }^{\circ}\text{C}$  and  $3.50\text{ atm}$ ?

Using the equation  $PV = nRT$ , assign the known and unknown variables. Typically, for single value problems this is easy. In this case;

$$V = \frac{nRT}{P}$$

$T = 25 + 273 = 298\text{ K}$ :  $V =$  is unknown:  $P = 3.50\text{ atm}$ :  $n =$  must be calculated.

$$n = 3.56\text{ g N}_2 \left( \frac{1\text{ mol}}{28.0\text{ g}} \right) = 0.127\text{ mol N}_2$$

$$(3.50\text{ atm})V = 0.127\text{ mol} (0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298\text{ K})$$

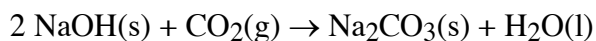
$$V = \frac{0.127\text{ mol} (0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298\text{ K})}{3.50\text{ atm}} = 0.888\text{ L}$$

4. Calculate the density of SF<sub>6</sub> at 1.00 atm and 0.00 °C.

$$\text{density} = \frac{P(\text{MW})}{RT}$$

$$\text{density} = \frac{1 \text{ atm} \left(146 \frac{\text{g}}{\text{mol}}\right)}{0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} (273 \text{ K})} = 6.51 \frac{\text{g}}{\text{L}}$$

5. Consider the reaction



which is a chemical means, although not economically viable, of removing CO<sub>2</sub> from the atmosphere. How many liters of CO<sub>2</sub> at 25.0 °C and 745 mmHg can be removed by 1.00 kg of NaOH?

$$1000 \text{ g NaOH} \left(\frac{1 \text{ mole}}{40 \text{ gm}}\right) = 25.0 \text{ mole NaOH}$$

$$25.0 \text{ mol NaOH} \left(\frac{1 \text{ mole CO}_2}{2 \text{ mole NaOH}}\right) = 12.5 \text{ mole CO}_2$$

$$V = \frac{n\text{CO}_2RT}{P} = \frac{12.5 \text{ moles} \cdot 0.0821 \left(\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right) (298 \text{ k})}{.980 \text{ atm}}$$

$$= 312 \text{ liters}$$

6. Calculate the total pressure in a 10.0 liter flask at 21 °C which contains 4.00 g H<sub>2</sub>, 12.0 g O<sub>2</sub> and 8.00 g He.

$$4.00 \text{ g H}_2 \left(\frac{1 \text{ mol}}{2.02 \text{ g}}\right) = 1.98 \text{ mol H}_2$$

$$12.00 \text{ g O}_2 \left(\frac{1 \text{ mol}}{32.0 \text{ g}}\right) = 0.375 \text{ mol O}_2$$

$$8.00 \text{ g He} \left(\frac{1 \text{ mol}}{4.0 \text{ g}}\right) = 2.00 \text{ mol He}$$

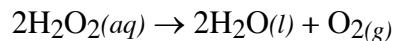
$$P_T = P_{\text{H}_2} + P_{\text{O}_2} + P_{\text{He}} = \frac{n_{\text{H}_2}RT}{V} + \frac{n_{\text{O}_2}RT}{V} + \frac{n_{\text{He}}RT}{V}$$

$$= (n_{\text{H}_2} + n_{\text{O}_2} + n_{\text{He}}) \left(\frac{RT}{V}\right)$$

$$P_T = (1.98 \text{ mol H}_2 + 0.375 \text{ mol O}_2 + 2.00 \text{ mol He}) \left(\frac{0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \cdot 294 \text{ K}}{10.0 \text{ L}}\right)$$

$$P_T = 10.5 \text{ atm}$$

7. A common laboratory preparation of O<sub>2</sub> involved the decomposition of hydrogen peroxide, H<sub>2</sub>O<sub>2</sub>, according to the equation,



If 240 mL of O<sub>2</sub> at 23 °C and at 0.965 atm pressure are collected over a sample of water at the same temperature determine the number of moles of O<sub>2</sub> obtained in the reaction.

$$P_T = P_{\text{O}_2} + P_{\text{H}_2\text{O}} \quad \text{at } 23^\circ\text{C} \text{ the } P_{\text{H}_2\text{O}} \text{ is } 21.0 \text{ mmHg}$$

$$P_{\text{O}_2} = P_T - P_{\text{H}_2\text{O}} = 733 \text{ mmHg} - 21.0 \text{ mmHg} = 712 \text{ mmHg} (.937 \text{ atm})$$

$$n_{\text{O}_2} = \frac{PV}{RT} = \frac{0.937 \text{ atm} \cdot 0.240 \text{ L}}{0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \cdot 296 \text{ K}} = 9.25 \times 10^{-3} \text{ mol O}_2$$