

A balloon is taken from a room where the temperature is 27.0 C to a freezer where the temperature is -5.0 C. If the balloon has a volume of 3.5 L in the 27.0 C room, what is the volume of the balloon in the freezer. Assume pressure is constant.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}, \text{ and } P_1 = P_2 \rightarrow \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{3.5 \text{ L}}{300.2 \text{ K}} = \frac{V_2}{268.2 \text{ K}}$$

$$\boxed{3.1 \text{ L} = V_2}$$

(volume in freezer)

$$\begin{aligned} V_1 &= 3.5 \text{ L} \\ T_1 &= 27.0^\circ\text{C} = 300.2 \text{ K} \\ V_2 &= ? \\ T_2 &= -5.0^\circ\text{C} = 268.2 \text{ K} \end{aligned}$$

2.25 L of nitrogen gas is trapped in a piston at 25.0 C and 1.00 atm pressure. If the piston is pushed in so that the gas's volume is 1.00 L while the temperature increases to 31.0 C, what is the pressure of the gas in the piston?

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$P_1 = 1.00 \text{ atm}$$

$$V_1 = 2.25 \text{ L}$$

$$T_1 = 25.0^\circ\text{C} = 298.2 \text{ K}$$

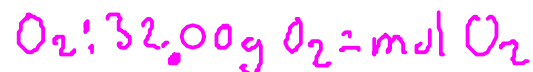
$$P_2 = ?$$

$$V_2 = 1.00 \text{ L}$$

$$T_2 = 31.0^\circ\text{C} = 304.2 \text{ K}$$

$$\frac{(1.00 \text{ atm})(2.25 \text{ L})}{(298.2 \text{ K})} = \frac{P_2 (1.00 \text{ L})}{(304.2 \text{ K})} ; P_2 = \boxed{2.30 \text{ atm}}$$

Calculate the mass of <sup>\*</sup>22650 L of oxygen gas at 25.0 C and 1.18 atm pressure.



\* Volume of a 10'x10'x8' room

- 1 - Use IDEAL GAS EQUATION ( $PV=nRT$ ) to find MOLES of oxygen gas.
- 2 - Convert moles oxygen gas to mass. Use FORMULA WEIGHT.

$$PV = nRT \quad \left| \quad \begin{array}{l} P = 1.18 \text{ atm} \\ V = 22650 \text{ L} \end{array} \quad \begin{array}{l} R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \\ T = 25.0^\circ \text{C} = 298.2 \text{ K} \end{array}$$

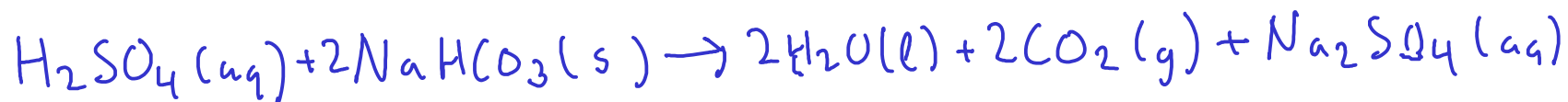
$$n = \frac{PV}{RT}$$

$$\textcircled{1} \quad n_{\text{O}_2} = \frac{(1.18 \text{ atm})(22650 \text{ L})}{(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298.2 \text{ K})} = 1092.222357 \text{ mol O}_2$$

$$\textcircled{2} \quad 1092.222357 \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{\text{mol O}_2} = \boxed{35000 \text{ g O}_2} \quad \begin{array}{l} 35.0 \text{ kg} \\ \sim 77 \text{ lb} \end{array}$$

## CHEMICAL CALCULATIONS WITH THE GAS LAWS

$$FW_{\text{NaHCO}_3} = 84.007 \text{ g/mol}$$



Given 25.0 g of sodium bicarbonate and sufficient sulfuric acid, what volume of carbon dioxide gas would be produced at 25.0 C and 0.950 atm pressure?

- 1 - Convert 25.0 grams sodium bicarbonate to moles. Use FORMULA WEIGHT.
- 2 - Convert moles sodium bicarbonate to moles carbon dioxide. Use CHEMICAL EQUATION
- 3 - Convert moles carbon dioxide to volume. Use IDEAL GAS EQUATION.

$$\textcircled{1} 84.007 \text{ g NaHCO}_3 = \text{mol NaHCO}_3 \quad \textcircled{2} 2 \text{ mol NaHCO}_3 = 2 \text{ mol CO}_2$$

$$25.0 \text{ g NaHCO}_3 \times \frac{\textcircled{1} \text{ mol NaHCO}_3}{84.007 \text{ g NaHCO}_3} \times \frac{\textcircled{2} 2 \text{ mol CO}_2}{2 \text{ mol NaHCO}_3} = 0.2975942481 \text{ mol CO}_2$$

$$\textcircled{3} \begin{array}{l} PV = nRT \\ V = \frac{nRT}{P} \end{array} \quad \begin{array}{l} n = 0.2975942481 \text{ mol CO}_2 \\ R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \\ T = 25.0^\circ\text{C} = 298.2 \text{ K} \\ P = 0.950 \text{ atm} \end{array}$$

$$V = \frac{(0.2975942481 \text{ mol CO}_2)(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298.2 \text{ K})}{(0.950 \text{ atm})} = 7.67 \text{ L} @ 25.0^\circ\text{C}, 0.950 \text{ atm}$$

What volume would the gas in the last example problem have at STP?

STP: "Standard Temperature and Pressure" (0 C and 1 atm)

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Use the combined gas law ...

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \left| \quad \begin{array}{l} P_1 = 0,950 \text{ atm} \\ V_1 = 7,67 \text{ L} \\ T_1 = 298,2 \text{ K} \end{array} \quad \begin{array}{l} P_2 = 1 \text{ atm} \\ V_2 = ? \\ T_2 = 0^\circ\text{C} = 273,2 \text{ K} \end{array} \right.$$

$$\frac{(0,950 \text{ atm})(7,67 \text{ L})}{(298,2 \text{ K})} = \frac{(1 \text{ atm})(V_2)}{(273,2 \text{ K})}$$

$$\boxed{\begin{array}{l} 6,67 \text{ L} \\ @ \text{ STP} \end{array}} = V_2$$



At  $300^\circ\text{C}$ , ammonium nitrate violently decomposes to produce nitrogen gas, oxygen gas, and water vapor. What is the total volume of gas that would be produced at 1.00 atm by the decomposition of 15.0 grams of ammonium nitrate?

To simplify the calculation, calculate the TOTAL MOLES OF GAS instead of treating each gas separately.

$$F_w \text{NH}_4\text{NO}_3 \approx 80.052 \text{ g/mol}$$

- 1 - Convert 15.0 grams ammonium nitrate to moles. Use FORMULA WEIGHT.
- 2 - Convert moles ammonium nitrate to TOTAL MOLES GAS. Use CHEMICAL EQUATION.
- 3 - Convert TOTAL MOLES OF GAS to volume. Use IDEAL GAS EQUATION.

$$\textcircled{1} 80.052 \text{ g NH}_4\text{NO}_3 = 1 \text{ mol NH}_4\text{NO}_3 \quad \textcircled{2} 2 \text{ mol NH}_4\text{NO}_3 = 7 \text{ mol gas} \quad (2+1+4=7)$$

$$15.0 \text{ g NH}_4\text{NO}_3 \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80.052 \text{ g NH}_4\text{NO}_3} \times \frac{7 \text{ mol gas}}{2 \text{ mol NH}_4\text{NO}_3} = 0.6558237146 \text{ mol gas}$$

$$\textcircled{3} \begin{array}{l} PV = nRT \\ V = \frac{nRT}{P} \end{array} \quad \begin{array}{l} n = 0.6558237146 \text{ mol gas} \quad P = 1.00 \text{ atm} \\ R = 0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \\ T = 300.^\circ\text{C} = 573 \text{ K} \end{array}$$

$$V = \frac{(0.6558237146 \text{ mol gas})(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(573 \text{ K})}{(1.00 \text{ atm})} = 30.8 \text{ L @ } 300.^\circ\text{C, } 1 \text{ atm}$$