

Gas Laws in Chemical Reactions

Problem 6-87

A mixture of 1.00 g H₂ and 8.60 g O₂ is introduced into a 1.500-L flask at 25°C. **(a)** What is the total gas pressure in the flask? **(b)** When the mixture is ignited, an explosive reaction occurs in which water is the only product. What is the total gas pressure when the flask is returned to 25°C? (The vapor pressure of water at 25°C is 23.8 mmHg.)

Solution:

This is a combination of a limiting reactant problem and a problem involving gas stoichiometry.

(a)

First, calculate molar quantities of the reactants:

$$n_{\text{H}_2} = 1.00 \text{ g H}_2 / 2.016 \frac{\text{g}}{\text{mol}} = 0.4960 \text{ mol H}_2$$

$$n_{\text{O}_2} = 8.60 \text{ g O}_2 / 32.00 \frac{\text{g}}{\text{mol}} = 0.2688 \text{ mol O}_2$$

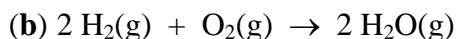
The pressure in the container is independent of the identities of the gases so calculate the total moles of gas and the pressure in the container:

$$n_{\text{total}} = 0.7648 \text{ mol}$$

$$T = 25^\circ\text{C} + 273.15 = 298.15 \text{ K}$$

$$V = 1.500 \text{ L}$$

$$P = \frac{nRT}{V} = \frac{(0.7648 \text{ mol})(0.082059 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298.15 \text{ K})}{1.500 \text{ L}} = 12.5 \text{ atm}$$



Now, work out the limiting reactant problem and calculate the amount of excess reactant:

Amount of H₂O produced:

$$n_{\text{H}_2} = 0.4960 \text{ mol H}_2 \quad n_{\text{H}_2\text{O}} = 0.4960 \text{ mol H}_2 \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} = 0.4960 \text{ mol H}_2\text{O}$$

$$n_{\text{O}_2} = 0.2688 \text{ mol O}_2 \quad n_{\text{H}_2\text{O}} = 0.2688 \text{ mol O}_2 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} = 0.5376 \text{ mol H}_2\text{O}$$

The H₂ is limiting

$$n_{\text{H}_2\text{O formed}} = 0.4960 \text{ mol H}_2\text{O}$$

$$n_{\text{O}_2 \text{ remaining}} = 0.2688 \text{ mol O}_2 - 0.4960 \text{ mol H}_2 \times \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2} = 0.0208 \text{ mol O}_2$$

When cooled to 25°C, almost all of the water vapor will condense to liquid water.

$$P = \frac{nRT}{V} = \frac{(0.0208 \text{ mol O}_2)(0.082059 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298.15 \text{ K})}{1.500 \text{ L}} = 0.339 \text{ atm}$$

However, there is a small amount of pressure contribution from the vapor pressure of the water (23.8 mmHg)

$$P_{\text{total}} = 0.339 \text{ atm} + 23.8 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.371 \text{ atm total pressure}$$