Gas Laws in Chemical Reactions

Problem 6-87

A mixture of 1.00 g H_2 and 8.60 g O_2 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:

$$n_{\mathrm{H_2}} = {}^{1.00\text{ g H}_2}\!\big/\!{}_{2.016\,\frac{\mathrm{g}}{\mathrm{mol}}} = 0.4960\text{ mol H}_2$$

$$n_{\mathrm{O_2}} = {}^{8.60\text{ g O}_2}\!\big/\!{}_{32.00\,\frac{\mathrm{g}}{\mathrm{mol}}} = 0.2688\text{ mol O}_2$$

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:

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

$$T = 25°\mathrm{C} + 273.15 = 298.15\text{ K}$$

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

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

(**b**) $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g)$

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

Amount of H_2O produced:

$$n_{\mathrm{H_2}} = 0.4960\text{ mol H}_2 \qquad n_{\mathrm{H_2O}} = 0.4960\text{ mol H}_2 \times \frac{2\text{ mol H}_2\text{O}}{2\text{ mol H}_2} = 0.4960\text{ mol H}_2\text{O}$$

$$n_{\mathrm{O_2}} = 0.2688\text{ mol O}_2 \qquad n_{\mathrm{H_2O}} = 0.2688\text{ mol O}_2 \times \frac{2\text{ mol H}_2\text{O}}{1\text{ mol H}_2} = 0.5376\text{ mol H}_2\text{O}$$

$\boxed{\text{The H}_2\text{ is limiting}}$

$$n_{H_2O\text{ formed}} = 0.4960\text{ mol } H_2O$$

$$n_{O_2\text{ remaining}} = 0.2688\text{ mol } O_2 - 0.4960\text{ mol } H_2 \times \frac{1\text{ mol } O_2}{2\text{ mol } H_2} = 0.0208\text{ mol } O_2$$

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

$$P = \frac{nRT}{V} = \frac{(0.0208\text{ mol } O_2)\left(0.082059\,\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298.15\text{ K})}{1.500\text{ L}} = 0.339\text{ atm}$$

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

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