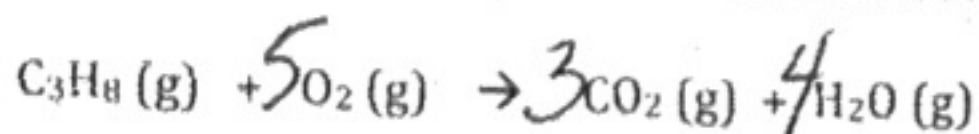


Below is an unbalanced chemical equation for the combustion of propane gas.



1. Balance the equation.

2. If you have 10L of  $\text{C}_3\text{H}_8$  gas in a container at 298K and a pressure of 1 atm, how many liters of  $\text{O}_2$  gas do you need to **completely** combust the propane? **VAN**

Given: 10 L  $\text{C}_3\text{H}_8$

Wanted: ? L  $\text{O}_2$

Conversion: 1 L  $\text{C}_3\text{H}_8$ : 5 L  $\text{O}_2$

$$\frac{10 \text{ L } \text{C}_3\text{H}_8}{1 \text{ L } \text{C}_3\text{H}_8} \times \frac{5 \text{ L } \text{O}_2}{1 \text{ L } \text{C}_3\text{H}_8} = \boxed{50 \text{ L } \text{O}_2}$$

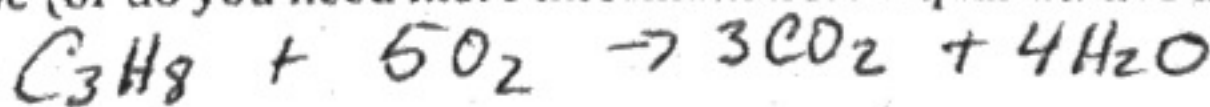
3. If you have a 20 L mixture of propane and oxygen gas with a mole fraction of propane,  $X_{\text{C}_3\text{H}_8} = 0.2$ , is there sufficient oxygen for complete combustion? Explain. **NO**

If  $X_{\text{C}_3\text{H}_8} = 0.2$ , then  $X_{\text{O}_2} = 0.8$  this is a 1:4 ratio

However from the balanced equation the ratio is 1:5

In general to avoid any side reaction you need an **excess of oxygen** for complete combustion. This is a very important fact, as people who combust fuels in a reduced oxygen environment can accidentally cause asphyxiation by the production of CO.

4. Assume you start with 4L of propane and 26 L of oxygen gas, both held a constant temperature and pressure. Assume that under these conditions, the reaction will be complete combustion (forming  $\text{CO}_2$  and  $\text{H}_2\text{O}$  as products). After the reaction will the volume of the container be the larger, smaller, or the same? What will the volume of the container be (or do you need more information for a quantitative answer)?  **$\text{C}_3\text{H}_8$  is limiting reagent**



$$\frac{4 \text{ L } \text{C}_3\text{H}_8}{1 \text{ L } \text{C}_3\text{H}_8} \times \frac{5 \text{ L } \text{O}_2}{1 \text{ L } \text{C}_3\text{H}_8} = 20 \text{ L } \text{O}_2 \text{ used}$$

$$\frac{4 \text{ L } \text{C}_3\text{H}_8}{1 \text{ L } \text{C}_3\text{H}_8} \times \frac{3 \text{ L } \text{CO}_2}{1 \text{ L } \text{C}_3\text{H}_8} = 12 \text{ L } \text{CO}_2 \text{ produced}$$

Now, assume that your combustion is **not complete** as the reaction will produce some carbon monoxide, CO, in addition to the  $\text{CO}_2$ . Writing and balancing an incomplete combustion is not trivial.

$$\frac{4 \text{ L } \text{C}_3\text{H}_8}{1 \text{ L } \text{C}_3\text{H}_8} \times \frac{4 \text{ L } \text{H}_2\text{O}}{1 \text{ L } \text{C}_3\text{H}_8} = 16 \text{ L } \text{H}_2\text{O} \text{ produced}$$

6 L  $\text{O}_2$  leftover  
 12 L  $\text{CO}_2$  produced  
 16 L  $\text{H}_2\text{O}$  produced

Initial Volume  $4\text{L} + 26\text{L} = 30\text{L}$

**LARGER** →

34 L Final Volume

5. Taking as the starting conditions 1 mole of propane and 5 moles of oxygen now held at constant volume of 20L and constant temperature of 300K. If the 1% of the total carbon forms CO, what is the partial pressure of CO after the reaction?

a) How much total C in rxn mixture?

$$1 \text{ mol } C_3H_8 = \boxed{3 \text{ mol } C}$$

b) How many moles CO formed?

$$3 \text{ mol } C \times 0.01 = \boxed{.03 \text{ mol } CO}$$

c) What is the partial pressure of CO in final mixture?

$$P_{CO} V_{CO} = n_{CO} R T_{CO}$$

$$V = 20 \text{ L}$$

$$n = .03 \text{ mol}$$

$$R = .08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

$$T = 300 \text{ K}$$

$$P_{CO} = \frac{nRT}{V} = \frac{.03 \cdot .08206 \cdot 300}{20}$$

$$\boxed{P_{CO} = .037 \text{ atm}}$$