9.5

Understand that the theoretical yield is rarely achieved in an experiment.

- **Theoretical Yield**
  - The maximum amount of product that could be produced during the chemical reaction.
- **Actual Yield**
  - The amount of product that is actually produced during a chemical reaction.

Sometimes, experiments don't go perfectly + you are left with less product than you should have.

- **Percent Yield**
  - $$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$
  - **Consider the following problem:** Suppose 6.85×10^4 g of carbon monoxide is reacted with 8.60×10^3 g of hydrogen gas to produce methanol (CH_3OH).
    - **(a) Calculate the theoretical yield of methanol. (b) If 3.57×10^4 g of methanol is actually produced, what is the percent yield of methanol?**
      - **(a)** $2H_2 + 1CO \rightarrow 1CH_3OH$

$$6.85 \times 10^4 \cancel{g\ CO} \times \frac{1 \text{ mol CO}}{28.01 \cancel{g\ CO}} = 2.45 \times 10^3 \text{ mol CO available}$$

$$8.60 \times 10^3 \cancel{g\ H_2} \times \frac{1 \text{ mol } H_2}{2.016 \cancel{g\ H_2}} = 4.27 \times 10^3 \text{ mol } H_2 \text{ available}$$

$$2.45 \times 10^3 \cancel{\text{mol CO}} \times \frac{2 \text{ mol } H_2}{1 \cancel{\text{mol CO}}} = 4.90 \times 10^3 \text{ mol } H_2 \text{ required}$$

H_2 is the limiting reactant because more is required then is available.

$$4.27 \times 10^3 \cancel{\text{mol } H_2} \times \frac{1 \cancel{\text{mol } CH_3OH}}{2 \cancel{\text{mol } H_2}} \times \frac{32.04 \text{ g } CH_3OH}{1 \cancel{\text{mol } CH_3OH}} = \boxed{6.86 \times 10^4 \text{ g } CH_3OH}$$

- **(b)**

$$\frac{3.57 \times 10^4 \text{ g}}{6.86 \times 10^4 \text{ g}} \times 100 = \boxed{52.0\%}$$

% yield should be below 100%.. If not, the answer is not reasonable.