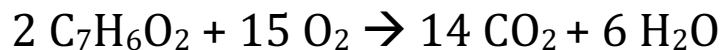


Station 1

The balanced equation for the combustion of benzoic acid is as follows:



A 305.0 g sample of $\text{C}_7\text{H}_6\text{O}_2$ is combined with 512.0 grams of O_2 .

a. Determine which reactant is in excess.

$$305.0 \text{g}_{\text{C}_7\text{H}_6\text{O}_2} \times \frac{1 \text{ mol}_{\text{C}_7\text{H}_6\text{O}_2}}{122.13 \text{g}_{\text{C}_7\text{H}_6\text{O}_2}} \times \frac{14 \text{ mol}_{\text{CO}_2}}{2 \text{ mol}_{\text{C}_7\text{H}_6\text{O}_2}} = 17.48 \text{ mol}_{\text{CO}_2} \leftarrow \text{excess reaction}$$

\therefore excess reactant = $\boxed{\text{C}_7\text{H}_6\text{O}_2}$

$$512.0 \text{g}_{\text{O}_2} \times \frac{1 \text{ mol}_{\text{O}_2}}{32.00 \text{g}_{\text{O}_2}} \times \frac{14 \text{ mol}_{\text{CO}_2}}{15 \text{ mol}_{\text{O}_2}} = 14.93 \text{ mol}_{\text{CO}_2} \leftarrow \text{limiting reaction}$$

b. When this reaction is carried out, what mass of CO_2 will be produced?

$$14.93 \text{ mol}_{\text{CO}_2} \times \frac{44.01 \text{g}_{\text{CO}_2}}{1 \text{ mol}_{\text{CO}_2}} = \boxed{657.1 \text{g}_{\text{CO}_2}}$$

c. Determine the mass of the excess reactant left over.

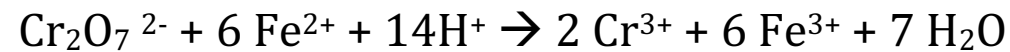
$$512.0 \text{g}_{\text{O}_2} \times \frac{1 \text{ mol}_{\text{O}_2}}{32.00 \text{g}_{\text{O}_2}} \times \frac{2 \text{ mol}_{\text{C}_7\text{H}_6\text{O}_2}}{15 \text{ mol}_{\text{O}_2}} \times \frac{122.13 \text{g}_{\text{C}_7\text{H}_6\text{O}_2}}{1 \text{ mol}_{\text{C}_7\text{H}_6\text{O}_2}} = 260.5 \text{g}_{\text{C}_7\text{H}_6\text{O}_2} \leftarrow \text{Used}$$

Have - Used = Excess

$$305.0 \text{g} - 260.5 \text{g} = \boxed{44.5 \text{g}_{\text{C}_7\text{H}_6\text{O}_2}} \quad * \text{subtraction sig figs}$$

Station 2

The iron present in a sample of iron ore is converted to Fe^{2+} and reacted with dichromate ion:



0.0176 L 0.0250 L

0.125 M

17.6 mL of 0.125 M dichromate is required to react with 25.0 mL sample of Fe^{2+} solution.

a. What is the molarity of Fe^{2+} ?

$$0.0176 \text{ L } \text{Cr}_2\text{O}_7^{2-} \times \frac{0.125 \text{ mol } \text{Cr}_2\text{O}_7^{2-}}{1 \text{ L } \text{Cr}_2\text{O}_7^{2-}} \times \frac{6 \text{ mol } \text{Fe}^{2+}}{1 \text{ mol } \text{Cr}_2\text{O}_7^{2-}} \times \frac{1}{0.0250 \text{ L } \text{Fe}^{2+}} = 0.528 \text{ M } \text{Fe}^{2+}$$

b. What mass of iron is present in the 25.0 mL sample?

$$0.0250 \text{ L } \text{Fe}^{2+} \times \frac{0.528 \text{ mol } \text{Fe}^{2+}}{1 \text{ L } \text{Fe}^{2+}} \times \frac{55.85 \text{ g } \text{Fe}^{2+}}{1 \text{ mol } \text{Fe}^{2+}} = 0.737 \text{ g } \text{Fe}^{2+}$$

Station 3

The reaction between nitrogen and hydrogen produces NH_3 .

a. What is the balanced equation?



b. At STP, calculate the volume of NH_3 that is produced when 145 L of N_2 reacts with excess hydrogen gas.

$$145 \text{ L N}_2 \times \frac{1 \text{ mol N}_2}{22.4 \text{ L N}_2} \times \frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2} \times \frac{22.4 \text{ L NH}_3}{1 \text{ mol NH}_3} = \boxed{290. \text{ L NH}_3}$$

c. How many litres of nitrogen react with 581 L of hydrogen at STP?

$$581 \text{ L H}_2 \times \frac{1 \text{ mol H}_2}{22.4 \text{ L H}_2} \times \frac{1 \text{ mol N}_2}{3 \text{ mol H}_2} \times \frac{22.4 \text{ L N}_2}{1 \text{ mol N}_2} = \boxed{194 \text{ L N}_2}$$

Station 4

Consider the following reaction:



a. What is the balanced equation?



b. If 6.01 g of Mg metal reacts with 8.45 g of HNO₃ at STP, what volume of H₂ gas is produced?

$$6.01 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{1 \text{ mol H}_2}{1 \text{ mol Mg}} \times \frac{22.4 \text{ L H}_2}{1 \text{ mol H}_2} = \cancel{5.54 \text{ L H}_2}$$

$$8.45 \text{ g HNO}_3 \times \frac{1 \text{ mol HNO}_3}{63.02 \text{ g HNO}_3} \times \frac{1 \text{ mol H}_2}{2 \text{ mol HNO}_3} \times \frac{22.4 \text{ L H}_2}{1 \text{ mol H}_2} = \boxed{1.50 \text{ L H}_2}$$

c. How much excess reactant is left over?

$$8.45 \text{ g HNO}_3 \times \frac{1 \text{ mol HNO}_3}{63.02 \text{ g HNO}_3} \times \frac{1 \text{ mol Mg}}{2 \text{ mol HNO}_3} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 1.63 \text{ g Mg}$$

↑ used

Have - Used = Excess

$$6.01 \text{ g} - 1.63 \text{ g} = \boxed{4.38 \text{ g Mg}}$$