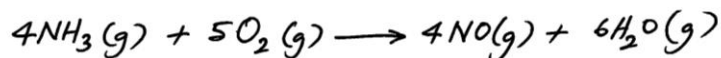


METHOD 3 : Example

One step in the industrial production of nitric acid is the reaction of ammonia with molecular oxygen to form nitrogen oxide.



In a study of this reaction, a chemist mixed 132g of ammonia with 273g oxygen and allowed them to react to completion. What is the limiting reactant.

Solution

We need the molar masses :

$$M_r(\text{NH}_3) = 17.03 \text{ g mol}^{-1}$$

$$M_r(\text{O}_2) = 32.00 \text{ g mol}^{-1}$$

$$M_r(\text{NO}) = 30.01 \text{ g mol}^{-1}$$

$$M_r(\text{H}_2\text{O}) = 18.02 \text{ g mol}^{-1}$$

Limiting Reactant

Convert masses to moles.

$$n(\text{moles}) = \frac{m(\text{grams})}{M_r(\text{molar mass, g mol}^{-1})}$$

$$n_{\text{NH}_3} = \frac{m_{\text{NH}_3}}{M_r(\text{NH}_3)} = \frac{132\text{g}}{17.03 \text{ g mol}^{-1}} = 7.75 \text{ mols.}$$

$$n_{\text{O}_2} = \frac{m_{\text{O}_2}}{M_r(\text{O}_2)} = \frac{273\text{g}}{32.00 \text{ g mol}^{-1}} = 8.53 \text{ mols.}$$

METHOD 3

- * The limiting reactant is the reactant that forms the lower mass of product / fewer moles of product.
- To identify NH_3 as the limiting reactant is to calculate the mass of NO that would be produced from 7.75 mols of NH_3 and excess of O_2

STOICHIOMETRY : LIMITING REACTANT : Method 3

OR • To identify O_2 as the limiting reactant is to calculate the mass of NO (the same product) that would be produced from 8.53 mols of O_2 and excess of NH_3 .

* The limiting reactant will be the one that produces the smaller mass/moles of NO.

If $NH_3 =$ limiting reactant (7.75 mols)

$$\frac{n_{NO}}{n_{NH_3}} = \frac{4}{4} \Rightarrow n_{NO} = n_{NH_3}$$

↑
stoichiometric coefficient

$$\frac{m_{NO}}{30.01 \text{ g mol}^{-1}} = 7.75 \text{ mols}$$
$$\therefore m_{NO} = (7.75 \times 30.01) \text{ g}$$
$$= \textcircled{232.6 \text{ g}} @ 7.75 \text{ mols}$$

If $O_2 =$ limiting reactant (8.53 mols)

$$\frac{n_{NO}}{n_{O_2}} = \frac{4}{5} \Rightarrow n_{NO} = \frac{4}{5} \times n_{O_2}$$

↑
stoichiometric coefficient

$$\frac{m_{NO}}{30.01} = \frac{4}{5} \times 8.53$$
$$\therefore m_{NO} = \left(\frac{4 \times 8.53 \times 30.01}{5} \right) \text{ g}$$
$$= \textcircled{204.8 \text{ g}} @ 6.82 \text{ mols}$$

This comparison shows that the amount of available oxygen (8.53 mols of O_2) would produce less NO than the amount of NH_3 available.

$\therefore O_2 =$ limiting reactant
 $NH_3 =$ excess reactant.