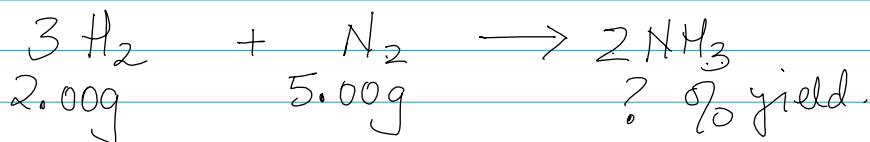


Limiting Reagent, % yield, Xs reagent.

Q) If 4.05g ammonia is produced from 2.00g of hydrogen and 5.00g of nitrogen; what is the percent yield of the reaction? ~~what~~ How much of the excess reagent is left over?



Ans Theoretical Yield calculation.

$$\begin{array}{l} \text{N}_2 \quad 5.00\text{g N}_2 \times \frac{1 \text{ mol N}_2}{28.00\text{g N}_2} \times \frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2} = \boxed{0.357 \text{ mol NH}_3} \quad \text{smaller.} \\ \text{H}_2 \quad 2.00\text{g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} = 0.661 \text{ mol NH}_3 \end{array}$$

$\times \frac{17.00 \text{ g NH}_3}{1 \text{ mol NH}_3} = \boxed{6.069 \text{ g}}$

Theoretical Yield = 6.069g

$$\% \text{ yield} = \frac{\text{AY}}{\text{TY}} \times 100\% = \frac{4.05\text{g}}{6.069\text{g}} \times 100\% = \boxed{66.7\%}$$

Excess Reagent (H₂)

$$\begin{array}{l} \text{mol NH}_3 \quad 0.661 \text{ mol NH}_3 - 0.357 \text{ mol NH}_3 = 0.304 \text{ mol NH}_3 \text{ excess.} \\ \text{mol NH}_3 \longrightarrow \text{mol H}_2 \longrightarrow \text{g H}_2 \\ \text{because H}_2 \text{ is in excess.} \end{array}$$

$$0.304 \text{ mol NH}_3 \times \frac{3 \text{ mol H}_2}{2 \text{ mol NH}_3} \times \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} = \boxed{0.919 \text{ g H}_2}$$

$$0.919 \text{ g H}_2 < 2.00 \text{ g H}_2 \text{ (cross check)}$$