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?

$$\underset{2.00g}{3H_2} + \underset{5.00g}{N_2} \longrightarrow \underset{?\ \%\ \text{yield}}{2NH_3}$$

Ans Theoretical Yield calculation:

N_2: $5.00g\,N_2 \times \frac{1\ mol\,N_2}{28.00g\,N_2} \times \frac{2\ mol\,NH_3}{1\ mol\,N_2} = \boxed{\underset{\text{smaller.}}{0.357 mol\ NH_3}} \times \frac{17.00\ g\,NH_3}{1\ mol\,NH_3} = \boxed{6.069g}$

H_2: $2.00g\,H_2 \times \frac{1\ mol\,H_2}{2.016\ g\,H_2} \times \frac{2\ mol\,NH_3}{3\ mol\,H_2} = 0.661 mol\ NH_3$

Theoretical Yield = 6.069g

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

Excess Reagent (H_2)

mol NH_3: $0.661\ mol\,NH_3 - 0.357 mol\,NH_3 = 0.304 mol\ NH_3$ excess.

$mol\,NH_3 \longrightarrow \underbrace{mol\,H_2}_{\text{because } H_2 \text{ is in excess.}} \longrightarrow g\,H_2$

$0.304 mol\,NH_3 \times \frac{3\ mol\,H_2}{2\ mol\,NH_3} \times \frac{2.016\ g\,H_2}{1\ mol\,H_2} = \boxed{0.919g\ H_2}$

$0.919g\,H_2 < 2.00g\,H_2$ (cross check)