Answer on Question #70867 - Chemistry - General Chemistry

Question:

Hydrogen and nitrogen gases are reacted to produce 2.10x10^3 kg of ammonia. if the percent yield of the reaction was only 20%, what mass of hydrogen was initially present. Assume hydrogen is the limiting reactant.

N2(g) + 3H2(g) = 2NH3(g)

Solution:

As one can see from the reaction, one mole of nitrogen reacts with 3 moles of hydrogen to produce 2 moles of ammonia:

$$\frac{n(N_2)}{1} = \frac{n(H_2)}{3} = \frac{n(NH_3)}{2}.$$

Then, knowing the number of the moles of the ammonia, we can calculate the quantity of hydrogen used.

Let's calculate the theoretical mass of ammonia produced, taking into account 20% yield:

$$m\left(NH_{3(theor)}\right) = m\left(NH_{3(exp)}\right) \cdot \left(\frac{100}{20}\right) = 2.10 \cdot 10^3 \cdot 5 = 1.05 \cdot 10^4 (kg).$$

Now it is easy to calculate the theoretical number of the moles of ammonia produced is:

$$n(NH_3) = \frac{m(NH_3)}{M(NH_3)} = \frac{1.05 \cdot 10^4 \cdot 10^3 (g)}{17.031\ (g\ mol^{-1})} = 6.17 \cdot 10^5 (mol).$$

Number of the moles of hydrogen initially present is:

$$n(H_2) = 3 \cdot \frac{n(NH_3)}{2} = 3 \cdot \frac{6.17 \cdot 10^5 (mol)}{2} = 9.25 \cdot 10^5 (mol).$$

And finally, the mass of hydrogen initially present:

$$m(H_2) = n(H_2) \cdot M(H_2) = 9.25 \cdot 10^5 (mol) \cdot 2.016\ (g\ mol^{-1}) = 1.86 \cdot 10^6 (g)$$

$$m(H_2) = 1.86 \cdot 10^3 (kg).$$

Answer: $1.86 \cdot 10^3 (kg)$