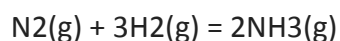


Answer on Question #70867 - Chemistry - General Chemistry

Question:

Hydrogen and nitrogen gases are reacted to produce 2.10×10^3 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.



Solution:

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

$$\frac{n(\text{N}_2)}{1} = \frac{n(\text{H}_2)}{3} = \frac{n(\text{NH}_3)}{2}.$$

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:

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

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

$$n(\text{NH}_3) = \frac{m(\text{NH}_3)}{M(\text{NH}_3)} = \frac{1.05 \cdot 10^4 \cdot 10^3 (\text{g})}{17.031 (\text{g mol}^{-1})} = 6.17 \cdot 10^5 (\text{mol}).$$

Number of the moles of hydrogen initially present is:

$$n(\text{H}_2) = 3 \cdot \frac{n(\text{NH}_3)}{2} = 3 \cdot \frac{6.17 \cdot 10^5 (\text{mol})}{2} = 9.25 \cdot 10^5 (\text{mol}).$$

And finally, the mass of hydrogen initially present:

$$m(\text{H}_2) = n(\text{H}_2) \cdot M(\text{H}_2) = 9.25 \cdot 10^5 (\text{mol}) \cdot 2.016 (\text{g mol}^{-1}) = 1.86 \cdot 10^6 (\text{g})$$

$$m(\text{H}_2) = 1.86 \cdot 10^3 (\text{kg}).$$

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

Answer provided by <https://www.AssignmentExpert.com>