Answer on Question#71929 – Chemistry – General chemistry

Question:

Magnesium nitride is formed in the reaction of magnesium metal with nitrogen gas in this reaction:

3 Mg(s) + N2(g) Mg3N2(s)

How many grams of product are formed from 2.0 mol of N2 (g) and 8.0 mol of Mg(s)?

Solution:

$$3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$$

1. Determine the limiting reagent:

$$\text{Actual ratio} = \frac{8.0 \text{ mol Mg(s)}}{2.0 \text{ mol } N_2} = \frac{4.0}{1.0}$$

$$\text{Stoichiometric ratio} = \frac{3 \text{ mol Mg(s)}}{1 \text{ mol } N_2(g)} = \frac{3}{1}$$

Actual ratio is greater than stoichiometric ratio, so N_2(g) is limiting reagent and Mg(s) is in excess.

2. Find how many moles of the product are formed:

$$n(Mg_3N_2) = n(N_2) = 2.0 \text{ mol}$$

3. Find how many mass of the product are formed:

$$m(Mg_3N_2) = n(Mg_3N_2) \times M(Mg_3N_2) = 2.0 \text{ mol} \times 100.95 \frac{g}{mol} = 201.9 \text{ g} = 2.0 \times 10^2 \text{g}$$

Explanation of significant digits:

Moles of N_2 are given with two significant digits (**2.0** mol of N_2 (g)), so the answer have to have two significant digits as well.

201.9 g has four significant digits

2.0×10^2g has two significant digits.

Answer:

$2.0 \times 10^2 \text{g } Mg_3N_2$