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CHAPTER 12**STUDY GUIDE FOR CONTENT MASTERY****Section 12.3 Limiting Reactants**

In your textbook, read about why reactions stop and how to determine the limiting reactant.

Study the diagram showing a chemical reaction and the chemical equation that represents the reaction. Then complete the table. Show your calculations for questions 15–17 in the space below the table.



The molar masses of O_2 , H_2 , and H_2O are 32.00 g/mol, 2.02 g/mol, and 18.02 g/mol, respectively.

Amount of O_2	Amount of H_2	Amount of H_2O	Limiting Reactant	Amount and Mass of Excess Reactant
1 molecule	2 molecules	2 molecules	none	none
4 molecules	4 molecules	4 molecules	H_2	2 molecules O_2
2 molecules	4 molecules	4 molecules	O_2	2 molecules H_2
1.00 mol	2.00 mol	2.00 mol	none	none
4.00 mol	4.00 mol	4.00 mol	H_2	2.00 mol O_2
2.00 mol	4.00 mol	4.00 mol	O_2	2.00 mol H_2
1.00 mol	4.00 mol	2.00 mol	O_2	2.00 mol H_2
4.000 mol	2.000 mol	2.000 mol	H_2	2.000 mol O_2
32.00 g	4.05 g	36.05 g	none	none
64.00 g	8.09 g	72.11 g	O_2	32.00 g H_2
64.00 g	20.00 g	36.05 g	O_2	32.00 g H_2

balanced equation mole ratio = 2 mol H_2O /1 mol O_2

$16.00 \text{ g } \text{O}_2 \times 1 \text{ mol } \text{O}_2 / 32.00 \text{ g } \text{O}_2 = 0.500 \text{ mol } \text{O}_2$

$20.00 \text{ g } \text{H}_2 \times 1 \text{ mol } \text{H}_2 / 2.02 \text{ g } \text{H}_2 = 9.901 \text{ mol } \text{H}_2$

actual mole ratio = $9.901 \text{ mol } \text{H}_2 / 0.500 \text{ mol } \text{O}_2 = 19.802 \text{ mol } \text{H}_2 / 1.000 \text{ mol } \text{O}_2$

Because the actual mole ratio of H_2/O_2 is larger than the balanced equation mole ratio of H_2/O_2 , there is an excess of H_2 ; O_2 is the limiting reactant.

Mass of H_2O used = $0.500 \text{ mol } \text{O}_2 \times 2 \text{ mol } \text{H}_2\text{O} / 1 \text{ mol } \text{O}_2 = 1.000 \text{ mol } \text{H}_2\text{O}$

$1.000 \text{ mol } \text{H}_2\text{O} \times 18.02 \text{ g } \text{H}_2\text{O} / 1 \text{ mol } \text{H}_2\text{O} = 18.02 \text{ g } \text{H}_2\text{O}$

Mass of H_2 produced = $0.500 \text{ mol } \text{H}_2\text{O} \times 2 \text{ mol } \text{H}_2 / 1 \text{ mol } \text{H}_2\text{O} = 1.000 \text{ mol } \text{H}_2$

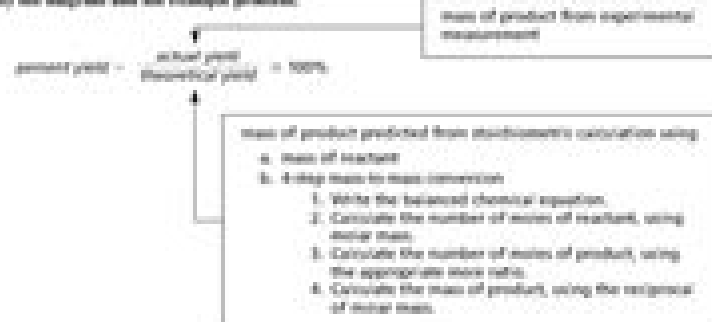
Excess H_2 = $9.901 \text{ mol } \text{H}_2 - 1.000 \text{ mol } \text{H}_2 = 8.901 \text{ mol } \text{H}_2$

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CHAPTER 13**STUDY GUIDE FOR CONTENT MASTERY****Section 13.4 Percent Yield**

In your textbook, read about the yields of products.

Study the diagram and the example problem.



Example Problem: The following chemical equation represents the production of palladium, a valuable metal in the manufacturing of some semiconductor devices.



In one experiment, the reaction yielded 7.42 g of the solid from a 7.00-g sample of palladium. Determine the percent yield of this reaction. The molar masses of Ga and $\text{Ga}_2(\text{NO}_3)_6$ are 69.72 g/mol and 137.44 g/mol, respectively.

Use the information in the diagram and example problem to evaluate each value or expression below. If the value or expression is correct, write correct. If it is incorrect, write the correct value or expression.

- actual yield: unknown 7.42 g $\text{Ga}_2(\text{NO}_3)_6$
- mass of reactant: 7.00 g Ga correct
- number of moles of reactant: $7.00 \text{ g } \text{Ga} \times \frac{69.72 \text{ g } \text{Ga}}{1 \text{ mol } \text{Ga}} = 1 \text{ mol } \text{Ga}$ 7.00 g $\text{Ga} \times 1 \text{ mol } \text{Ga} / 69.72 \text{ g } \text{Ga}$
- number of moles of product: $1 \text{ mol } \text{Ga} \times \frac{2 \text{ mol } \text{Ga}_2(\text{NO}_3)_6}{1 \text{ mol } \text{Ga}} = 2 \text{ mol } \text{Ga}_2(\text{NO}_3)_6$ 2.00 mol $\text{Ga}_2(\text{NO}_3)_6$
- theoretical yield: $1 \text{ mol } \text{Ga}_2(\text{NO}_3)_6 \times \frac{137.44 \text{ g } \text{Ga}_2(\text{NO}_3)_6}{1 \text{ mol } \text{Ga}_2(\text{NO}_3)_6} = 137.44 \text{ g}$ correct
- percent yield: $\frac{7.42 \text{ g } \text{Ga}_2(\text{NO}_3)_6}{137.44 \text{ g } \text{Ga}_2(\text{NO}_3)_6} \times 100 = 5.39\%$ 7.42 g $\text{Ga}_2(\text{NO}_3)_6 / 137.44 \text{ g } \text{Ga}_2(\text{NO}_3)_6 \times 100$

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