

Name _____ Date _____ **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_2O , and H_2 are 32.00 g/mol, 18.02 g/mol, and 2.02 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
2 molecules	4 molecules	4 molecules	H_2	2 molecules O_2
3 molecules	6 molecules	6 molecules	O_2	3 molecules H_2
1.00 mol	2.00 mol	2.00 mol	H_2	none
2.00 mol	4.00 mol	4.00 mol	H_2	2.00 mol O_2
3.00 mol	6.00 mol	6.00 mol	O_2	3.00 mol H_2
1.00 mol	2.00 mol	2.00 mol	H_2	none
2.00 mol	4.00 mol	4.00 mol	H_2	2.00 mol O_2
3.00 mol	6.00 mol	6.00 mol	O_2	3.00 mol H_2
1.00 mol	2.00 mol	2.00 mol	H_2	none
2.00 mol	4.00 mol	4.00 mol	H_2	2.00 mol O_2
3.00 mol	6.00 mol	6.00 mol	O_2	3.00 mol H_2
1.00 g	2.00 g	2.00 g	H_2	none
2.00 g	4.00 g	4.00 g	H_2	2.00 g O_2
3.00 g	6.00 g	6.00 g	O_2	3.00 g H_2

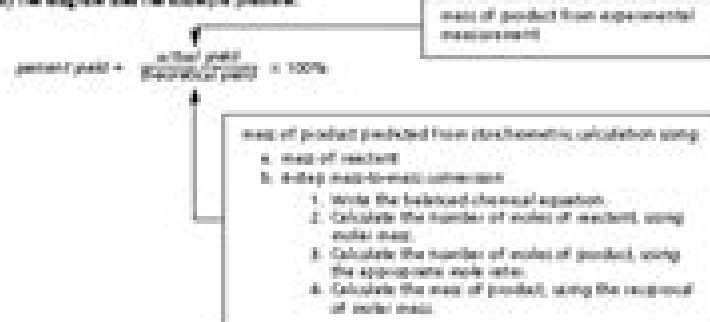
balanced equation mole ratio = 2 mol H_2O /1 mol O_2
 $16.00 \text{ g } \text{O}_2 \times 1 \text{ mol } \text{O}_2 / 32.00 \text{ g } \text{O}_2 = 0.500 \text{ mol } \text{O}_2$
 $20.00 \text{ g } \text{H}_2 \times 1 \text{ mol } \text{H}_2 / 2.02 \text{ g } \text{H}_2 = 9.90 \text{ mol } \text{H}_2$
 actual mole ratio = 9.90 mol H_2 /0.500 mol O_2 = 19.8 mol H_2 /1.000 mol 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 mol $\text{O}_2 \times 2 \text{ mol } \text{H}_2\text{O} / 1 \text{ mol } \text{O}_2 = 1.00 \text{ mol } \text{H}_2\text{O}$
 $1.00 \text{ mol } \text{H}_2\text{O} \times 18.02 \text{ g } \text{H}_2\text{O} / 1 \text{ mol } \text{H}_2\text{O} = 18.02 \text{ g } \text{H}_2\text{O}$
 Mass of H_2O produced = 0.500 mol $\text{O}_2 \times 2 \text{ mol } \text{H}_2\text{O} / 1 \text{ mol } \text{O}_2 = 1.00 \text{ mol } \text{H}_2\text{O}$
 $1.00 \text{ mol } \text{H}_2\text{O} \times 18.02 \text{ g } \text{H}_2\text{O} / 1 \text{ mol } \text{H}_2\text{O} = 18.02 \text{ g } \text{H}_2\text{O}$
 Excess H_2 = 20.00 g H_2 - 18.02 g H_2 = 1.98 g H_2

Name _____ Date _____ **CHAPTER 12** **STUDY GUIDE FOR CONTENT MASTERY**

Section 12.4 Percent Yield

In your textbook, read about the yield of products.

Study the diagram and the example problem.



Example Problem: The following chemical equation represents the production of gallium nitride, a substance used in the manufacturing of some semiconductor devices.



In one experiment, the reaction yielded 7.42 g of the oxide from a 1.00-g sample of gallium. Determine the percent yield of this reaction. The molar masses of Ga and Ga_3N_2 are 69.72 g/mol and 177.48 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.

1. actual yield, substance = 7.42 g Ga_3N_2
2. mass of reactant = 1.00 g Ga
3. number of moles of reactant = $1.00 \text{ g Ga} \times \frac{1 \text{ mol Ga}}{69.72 \text{ g Ga}} = 0.0143 \text{ mol Ga}$
4. number of moles of product = $0.0143 \text{ mol Ga} \times \frac{2 \text{ mol Ga}_3\text{N}_2}{3 \text{ mol Ga}} = 0.00953 \text{ mol Ga}_3\text{N}_2$
5. theoretical yield = $0.00953 \text{ mol Ga}_3\text{N}_2 \times \frac{177.48 \text{ g Ga}_3\text{N}_2}{1 \text{ mol Ga}_3\text{N}_2} = 1.69 \text{ g Ga}_3\text{N}_2$
6. percent yield = $\frac{7.42 \text{ g Ga}_3\text{N}_2}{1.69 \text{ g Ga}_3\text{N}_2} \times 100 = 439\%$

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James Michael Stewart



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