

Answers Questions 9.2

1. Complete the table below for the equation  $2\text{Al} + 3\text{S} \rightarrow \text{Al}_2\text{S}_3$

| Element or Compound       | Al                             | S                              | $\text{Al}_2\text{S}_3$ |
|---------------------------|--------------------------------|--------------------------------|-------------------------|
| Number of atoms/molecules | $2 \times 6.02 \times 10^{23}$ | $3 \times 6.02 \times 10^{23}$ | $6.02 \times 10^{23}$   |
| Number of moles           | 2.0                            | 3.0                            | 1.0                     |
| Number of grams           | 54.0                           | 96.0                           | 150.0                   |

2. Answer the questions below about the reaction  $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$

A. If 5.00 moles of hydrogen gas were reacted with 2.00 moles of nitrogen gas how many moles of ammonia would be produced?

$$5.00 \text{ moles H}_2 * \frac{2 \text{ moles NH}_3}{3 \text{ moles H}_2} = 3.33 \text{ moles NH}_3$$

$$2.00 \text{ moles N}_2 * \frac{2 \text{ moles NH}_3}{1 \text{ mole N}_2} = 4.00 \text{ moles NH}_3$$

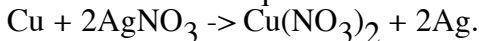
Would the reaction run out of hydrogen gas or nitrogen gas first?

Hydrogen gas runs out first, producing a smaller amount of ammonia

B. How many moles of nitrogen gas would be needed to react completely with 1.50 moles of hydrogen gas with no nitrogen gas left over?

$$1.5 \text{ moles H}_2 * \frac{1 \text{ mole N}_2}{3 \text{ moles H}_2} = 0.5 \text{ mole N}_2$$

3. Answer the questions below about the reaction



A. If 63.5 grams of copper metal reacted with 231.9 grams of silver nitrate, what mass of silver metal would be produced? Which reactant would run out first?

$$63.5 \text{ g Cu} * \frac{1 \text{ mole Cu} * 2 \text{ mole Ag} * 107.9 \text{ g Ag}}{63.5 \text{ g Cu} * 1 \text{ mole Cu} * 1 \text{ mole Ag}} = 215.8 \text{ g Ag}$$

$$231.9 \text{ g AgNO}_3 * \frac{1 \text{ mole AgNO}_3 * 2 \text{ mole Ag} * 107.9 \text{ g Ag}}{169.9 \text{ g AgNO}_3 * 2 \text{ mole AgNO}_3 * 1 \text{ mole Ag}} = 148 \text{ g Ag}$$

B. If 107.9 grams of silver were produced in a reaction, how many grams of cupric nitrate would be produced?

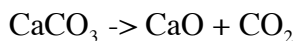
$$107.9 \text{ g Ag} * \frac{1 \text{ mole Ag} * 1 \text{ mole Cu}(\text{NO}_3)_2 * 187.5 \text{ g Cu}(\text{NO}_3)_2}{107.9 \text{ g Ag} * 2 \text{ mole Ag} * 1 \text{ mole Cu}(\text{NO}_3)_2} = 93.8 \text{ g Cu}(\text{NO}_3)_2$$

C. If you wanted to produce 10.79 grams of silver metal, what minimum number of grams of copper metal and silver nitrate would you have to start the reaction with?

$$10.79 \text{ g Ag} * \frac{1 \text{ mole Ag} * 1 \text{ mole Cu} * 63.5 \text{ g Cu}}{107.9 \text{ g Ag} * 2 \text{ mole Ag} * 1 \text{ mole Cu}} = 3.175 \text{ g Cu}$$

$$10.79 \text{ g Ag} * \frac{1 \text{ mole Ag} * 1 \text{ mole AgNO}_3 * 169.9 \text{ g AgNO}_3}{107.9 \text{ g Ag} * 1 \text{ mole Ag} * 1 \text{ mole AgNO}_3} = 16.99 \text{ g AgNO}_3$$

4. 275 grams of calcium carbonate is heated to produce calcium oxide and carbon dioxide gas. What mass of carbon dioxide gas and calcium oxide are produced?



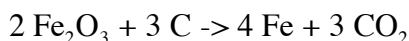
$$275\text{g CaCO}_3 * \frac{1 \text{ mole CaCO}_3}{100.1\text{g CaCO}_3} * \frac{1 \text{ mole CaO}}{1 \text{ mole CaCO}_3} * \frac{56.1\text{g CaO}}{1 \text{ mole CaO}} = 154\text{g CaO}$$

Total mass of reactants = total mass of products

$$\text{Mass CO}_2 = \text{mass CaCO}_3 - \text{mass CaO} = 275\text{g} - 154\text{g} = 121\text{g CO}_2$$

5. Ferric oxide can be converted to iron metal and carbon dioxide gas by heating with carbon.

A. What mass of carbon is needed to completely convert 385 grams of ferric oxide ?



$$385\text{g Fe}_2\text{O}_3 * \frac{1 \text{ mole Fe}_2\text{O}_3}{159.6\text{g Fe}_2\text{O}_3} * \frac{3 \text{ mole C}}{2 \text{ mole Fe}_2\text{O}_3} * \frac{12.0\text{g C}}{1 \text{ mole C}} = 43.4\text{g C}$$

B. How many grams of iron metal are produced?

$$385\text{g Fe}_2\text{O}_3 * \frac{1 \text{ mole Fe}_2\text{O}_3}{159.6\text{g Fe}_2\text{O}_3} * \frac{4 \text{ mole Fe}}{2 \text{ mole Fe}_2\text{O}_3} * \frac{55.8\text{g Fe}}{1 \text{ mole Fe}} = 269\text{g Fe}$$

6. Why is the theoretical yield of a chemical reaction is always greater than the actual yield ?

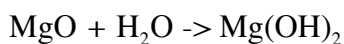
The reactants may not be pure

The reaction may not occur completely

Some loss of product can occur in the lab process

7. 80.6 grams of magnesium oxide is reacted with unlimited water to produce magnesium hydroxide.

A. Write and balance the equation for the reaction.



B. How many grams of magnesium hydroxide should be produced ? (show your work)

$$80.6\text{g MgO} * \frac{1 \text{ mole MgO}}{40.3\text{g MgO}} * \frac{1 \text{ mole Mg(OH)}_2}{1 \text{ mole MgO}} * \frac{58.3\text{g Mg(OH)}_2}{1 \text{ mole Mg(OH)}_2}$$

$$= 117\text{g Mg(OH)}_2$$

C. If 105 grams of magnesium hydroxide is actually produced, what is the percent yield for the reaction ? (show your work)

$$\% \text{ yield} = \frac{\text{actual mass in lab}}{\text{calculated mass expected}} = \frac{105\text{g}}{117\text{g}} = 89.7\%$$