Basic Chemistry 6th Edition
Zumdahl & DeCoste 2008

Chapter 9 Review

Key Terms

mole ratio (9.2)
stoichiometry (9.3)
limiting reactant (limiting reagent) (9.4)
theoretical yield (9.5)
percent yield (9.5)

Summary

1. A balanced equation relates the numbers of molecules of reactants and products. It can also be expressed in terms of the numbers of moles of reactants and products.
2. The process of using a chemical equation to calculate the relative amounts of reactants and products involved in the reaction is called doing stoichiometric calculations. To convert between moles of reactants and moles of products, we use mole ratios derived from the balanced equation.
3. Often reactants are not mixed in stoichiometric quantities (they do not "run out" at the same time). In that case, we must use the limiting reactant to calculate the amounts of products formed.
4. The actual yield of a reaction is usually less than its theoretical yield. The actual yield is often expressed as a percentage of the theoretical yield, which is called the percent yield.

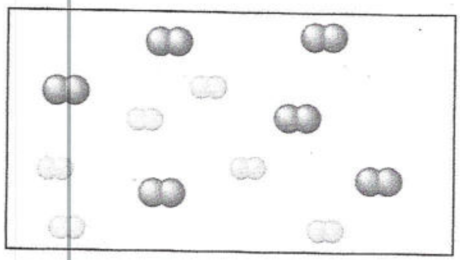
Active Learning Questions

These questions are designed to be considered by groups of students in class. Often these questions work well for introducing a particular topic in class.

1. Relate Active Learning Question 2 from Chapter 2 to the concepts of chemical stoichiometry.
2. You are making cookies and are missing a key ingredient—eggs. You have plenty of the other ingredients, except that you have only 1.33 cups of butter and no eggs. You note that the recipe calls for 2 cups of butter and 3 eggs (plus the other ingredients) to make 6 dozen cookies. You telephone a friend and have him bring you some eggs.
   a. How many eggs do you need?
   b. If you use all the butter (and get enough eggs), how many cookies can you make?

   Unfortunately, your friend hangs up before you tell him how many eggs you need. When he arrives, he has a surprise for you—to save time he has broken the eggs in a bowl for you. You ask him how many he brought, and he replies, "All of them, but I spilled some on the way over." You weigh the eggs and find that they weigh 62.1 g. Assuming that an average egg weighs 34.21 g:

   c. How much butter is needed to react with all the eggs?
   d. How many cookies can you make?
   e. Which will you have left over, eggs or butter?
   f. How much is left over?
   g. Relate this question to the concepts of chemical stoichiometry.
3. Nitrogen (N_2) and hydrogen (H_2) react to form ammonia (NH_3). Consider the mixture of N_2 ( ) and H_2( ) in a closed container as illustrated below:



   Assuming the reaction goes to completion, draw a representation of the product mixture. Explain how you arrived at this representation.
4. Which of the following equations best represents the reaction for Question 3?
   a. $6N_2 + 6H_2 \rightarrow 4NH_3 + 4N_2$
   b. $N_2 + H_2 \rightarrow NH_3$
   c. $N + 3H \rightarrow NH_3$
   d. $N_2 + 3H_2 \rightarrow 2NH_3$
   e. $2N_2 + 6H_2 \rightarrow 4NH_3$

   For choices you did not pick, explain what you feel is wrong with them, and justify the choice you did pick.
5. You know that chemical A reacts with chemical B. You react 10.0 g A with 10.0 g B. What information do you need to know to determine the amount of product that will be produced? Explain.
6. If 10.0 g of hydrogen gas is reacted with 10.0 g of oxygen gas, what mass of water can be produced?

Questions 7 and 8 deal with the following situation: You react chemical A with chemical B to make one product. It takes 100 g A to react completely with 20 g B.

7. What is the mass of the product?
   a. Less than 20 g
   b. Between 20 g and 100 g
   c. Between 100 g and 120 g
   d. Exactly 120 g
   e. More than 120 g

8. What is true about the chemical properties of the product?

   a. The properties are more like those of chemical A.
   b. The properties are more like those of chemical B.
   c. The properties are equally like those of chemical A and chemical B.
   d. The properties are not necessarily like either those of A or B.
   e. The properties are more like those of A or more like those of B, but more information is needed.

   For choices you did not pick, explain what you feel is wrong with them, and justify the choice you did pick.

9. The limiting reactant in a reaction:

   a. has the lowest coefficient in a balanced equation.
   b. is the reactant for which you have the fewest number of moles.
   c. has the lowest ratio: moles available/coefficient in the balanced equation.
   d. has the lowest ratio: coefficient in the balanced equation/moles available.
   e. None of the above.

   For choices you did not pick, explain what you feel is wrong with them, and justify the choice you did pick.

10. Given the equation $3A + B \rightarrow C + D$, if 4 moles of A is reacted with 2 moles of B, which of the following is true?

    a. The limiting reactant is the one with the higher molar mass.
    b. A is the limiting reactant because you need 6 moles of A and have 4 moles.
    c. B is the limiting reactant because you have fewer moles of B than moles of A.
    d. B is the limiting reactant because three A molecules react with every one B molecule.
    e. Neither reactant is limiting.

    For choices you did not pick, explain what you feel is wrong with them, and justify the choice you did pick.

11. A kerosene lamp has a mass of 1.5 kg. You put 0.5 kg of kerosene in the lamp. You burn all of the kerosene until the lamp has a mass of 1.5 kg. What is the mass of the gases given off? Explain.

12. What happens to the weight of an iron bar when it rusts?

    a. There is no change because mass is always conserved.
    b. The weight increases.
    c. The weight increases, but if the rust is scraped off, the bar has the original weight.
    d. The weight decreases.

    Justify your choice and, for choices you did not pick, explain what is wrong with them. Explain what it means for something to rust.

13. You may have noticed that water sometimes drips from the exhaust pipe of a car as it is running. Is this evidence that at least a small amount of water is originally present in the gasoline? Explain.

14. You have a chemical in a sealed glass container filled with air. The system has a mass of 250.0 g. The chemical is ignited by means of a magnifying glass focusing sunlight on the reactant. After the chemical is completely burned, what is the mass of the system? Explain your answer.

15. Consider the equation $2A + B \rightarrow A_2B$. If you mix 1.0 mol of A and 1.0 mol of B, how many moles of A_2B can be produced?

16. Can the percent yield of a reaction ever be greater than 100%? Explain.

17. According to the law of conservation of mass, mass cannot be gained or destroyed in a chemical reaction. So why can't you simply add the masses of two reactants to determine the total mass of product produced?

Questions and Problems

All even-numbered exercises have answers in the back of this book and solutions in the *Solutions Guide.*

9.1 Information Given by Chemical Equations

QUESTIONS

1. What do the coefficients of a balanced chemical equation tell us about the proportions in which atoms and molecules react on an individual (microscopic) basis?
2. What do the coefficients of a balanced chemical equation tell us about the proportions in which substances react on a macroscopic (mole) basis?
3. Although *mass* is a property of matter we can conveniently measure in the laboratory, the coefficients of a balanced chemical equation are *not* directly interpreted on the basis of mass. Explain why.
4. For the balanced chemical equation $H_2 + Br_2 \rightarrow 2HBr$, explain why we do *not* expect to produce 2 g of HBr if 1 g of H_2 is reacted with 1 g of Br_2.

PROBLEMS

5. For each of the following reactions, give the balanced equation for the reaction and state the meaning of the equation in terms of the numbers of *individual molecules* and in terms of *moles of molecules*.

   a. $PCl_3(l) + H_2O(l) \rightarrow H_3PO_3(aq) + HCl(g)$
   b. $XeF_2(g) + H_2O(l) \rightarrow Xe(g) + HF(g) + O_2(g)$
   c. $S(s) + HNO_3(aq) \rightarrow H_2SO_4(aq) + H_2O(l) + NO_2(g)$
   d. $NaHSO_3(s) \rightarrow Na_2SO_3(s) + SO_2(g) + H_2O(l)$

6. For each of the following reactions, balance the chemical equation and state the stoichiometric meaning of the equation in terms of the numbers of individual molecules reacting and in terms of *moles of molecules* reacting.

a. $(NH_4)_2CO_3(s) \rightarrow NH_3(g) + CO_2(g) + H_2O(g)$
b. $Mg(s) + P_4(s) \rightarrow Mg_3P_2(s)$
c. $Si(s) + S_8(s) \rightarrow Si_2S_4(l)$
d. $C_2H_5OH(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

9.2 Mole–Mole Relationships

QUESTIONS

7. Consider the reaction represented by the chemical equation

$$KOH(s) + SO_2(g) \rightarrow KHSO_3(s)$$

Since the coefficients of the balanced chemical equation are all equal to 1, we know that exactly 1 g of KOH will react with exactly 1 g of SO_2. True or false? Explain.

8. For the balanced chemical equation for the decomposition of hydrogen peroxide

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

explain why we know that decomposition of 2 g of hydrogen peroxide will *not* result in the production of 2 g of water and 1 g of oxygen gas.

9. Consider the balanced equation

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

What is the mole ratio that would enable you to calculate the number of moles of oxygen needed to react exactly with a given number of moles of $CH_4(g)$? What mole ratios would you use to calculate how many moles of each product form from a given number of moles of CH_4?

10. Consider the unbalanced chemical equation

$$S(s) + H_2SO_4(l) \rightarrow SO_2(g) + H_2O(l)$$

Balance the equation, and then write the mole ratios that would allow you to calculate the number of moles of each product that would form for a given number of moles of sulfur reacting. Write also the mole ratio that would allow you to calculate the number of moles of sulfuric acid that would be required to react with a given number of moles of sulfur.

PROBLEMS

11. For each of the following balanced reactions, calculate how many *moles of product* would be produced by complete conversion of 0.15 mol of the reactant indicated in boldface. State clearly the mole ratio used for the conversion.

a. $\mathbf{2Mg}(s) + O_2(g) \rightarrow 2MgO(s)$
b. $2Mg(s) + \mathbf{O_2}(g) \rightarrow 2MgO(s)$
c. $\mathbf{4Fe}(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$
d. $4Fe(s) + \mathbf{3O_2}(g) \rightarrow 2Fe_2O_3(s)$

12. For each of the following unbalanced equations, calculate how many *moles* of the *second reactant* would be required to react completely with *5.00 mol* of the *first reactant.*

a. $C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
b. $P_4(s) + O_2(g) \rightarrow P_4O_{10}(g)$
c. $CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$
d. $Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$

13. For each of the following balanced reactions, calculate how many *moles of each product* would be produced by complete conversion of 1.25 mol of the reactant indicated in boldface. State clearly the mole ratio used for the conversion.

a. $\mathbf{C_2H_5OH}(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$
b. $\mathbf{N_2}(g) + O_2(g) \rightarrow 2NO(g)$
c. $\mathbf{2NaClO_2}(s) + Cl_2(g) \rightarrow 2ClO_2(g) + 2NaCl(s)$
d. $\mathbf{3H_2}(g) + N_2(g) \rightarrow 2NH_3(g)$

14. For each of the following balanced chemical equations, calculate how many *moles* and how many *grams* of each product would be produced by the complete conversion of 0.50 mol of the reactant indicated in boldface. State clearly the mole ratio used for each conversion.

a. $\mathbf{NH_3}(g) + HCl(g) \rightarrow NH_4Cl(s)$
b. $CH_4(g) + \mathbf{4S}(s) \rightarrow CS_2(l) + 2H_2S(g)$
c. $\mathbf{PCl_3} + 3H_2O(l) \rightarrow H_3PO_3(aq) + 3HCl(aq)$
d. $\mathbf{NaOH}(s) + CO_2(g) \rightarrow NaHCO_3(s)$

15. For each of the following *unbalanced* equations, indicate how many *moles* of the *second reactant* would be required to react exactly with *0.275 mol* of the *first reactant.* State clearly the mole ratio used for the conversion.

a. $Cl_2(g) + KI(aq) \rightarrow I_2(s) + KCl(aq)$
b. $Co(s) + P_4(s) \rightarrow Co_3P_2(s)$
c. $Zn(s) + HNO_3(aq) \rightarrow ZnNO_3(aq) + H_2(g)$
d. $C_5H_{12}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

16. For each of the following *unbalanced* equations, indicate how many *moles* of the *first product* are produced if *0.625 mol* of the *second product* forms. State clearly the mole ratio used for each conversion.

a. $KO_2(s) + H_2O(l) \rightarrow O_2(g) + KOH(s)$
b. $SeO_2(g) + H_2Se(g) \rightarrow Se(s) + H_2O(g)$
c. $CH_3CH_2OH(l) + O_2(g) \rightarrow CH_3CHO(aq) + H_2O(l)$
d. $Fe_2O_3(s) + Al(s) \rightarrow Fe(l) + Al_2O_3(s)$

9.3 Mass Calculations

QUESTIONS

17. What quantity serves as the conversion factor between the mass of a sample and how many moles the sample contains?

18. What does it mean to say that the balanced chemical equation for a reaction describes the *stoichiometry* of the reaction?

PROBLEMS

19. Using the average atomic masses given inside the front cover of this book, calculate how many *moles* of each substance the following masses represent.

a. 2.62 g of helium gas, He
b. 4.95 g of boric acid, H_3BO_3

c. 8.31 g of calcium fluoride, CaF_2
d. 0.195 g of magnesium acetate, $Mg(C_2H_3O_2)_2$
e. 9.72 g of ammonia, NH_3

20. Using the average atomic masses given inside the front cover of this book, calculate how many *moles* of each substance the following masses represent.
a. 2.36 mg of lithium carbonate, Li_2CO_3
b. 1.92×10^{-3} g of uranium, U
c. 3.21 kg of lead chloride, $PbCl_2$
d. 4.62 g of glucose, $C_6H_{12}O_6$
e. 7.75 g of potassium hydroxide, KOH

21. Using the average atomic masses given inside the front cover of this book, calculate the *mass in grams* of each of the following samples.
a. 4.25 mol of oxygen gas, O_2
b. 1.27 millimol of platinum (1 millimol = 1/1000 mol)
c. 0.00101 mol of iron(II) sulfate, $FeSO_4$
d. 75.1 mol of calcium carbonate, $CaCO_3$
e. 1.35×10^{-4} mol of gold
f. 1.29 mol of hydrogen peroxide, H_2O_2
g. 6.14 mol of copper(II) sulfide, CuS

22. Using the average atomic masses given inside the front cover of this book, calculate the *mass in grams* of each of the following samples.
a. 0.624 mol of copper(I) iodide, CuI
b. 4.24 mol of bromine, Br_2
c. 0.000211 mol of xenon tetrafluoride, XeF_4
d. 9.11 mol of ethylene, C_2H_4
e. 1.21 millimol of ammonia, NH_3 (1 millimol = 1/1000 mol)
f. 4.25 mol of sodium hydroxide, NaOH
g. 1.27×10^{-6} mol of potassium iodide, KI

23. For each of the following *unbalanced* equations, calculate the *mass of each product* that could be produced by complete reaction of 1.55 g of the reactant indicated in boldface.
a. $\mathbf{CS_2}(l) + O_2(g) \rightarrow CO_2(g) + SO_2(g)$
b. $\mathbf{NaNO_3}(s) \rightarrow NaNO_2(s) + O_2(g)$
c. $H_2(g) + \mathbf{MnO_2}(s) \rightarrow MnO(s) + H_2O(g)$
d. $\mathbf{Br_2}(l) + Cl_2(g) \rightarrow BrCl(g)$

24. For each of the following *unbalanced* equations, calculate how many *moles* of the *second reactant* would be required to react completely with exactly 5.00 g of the *first reactant*.
a. $C_2H_5OH(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
b. $P_4(s) + O_2(g) \rightarrow P_4O_{10}(g)$
c. $MgO(s) + CO_2(g) \rightarrow MgCO_3(s)$
d. $Fe(s) + O_2(g) \rightarrow FeO(s)$

25. For each of the following *unbalanced* equations, calculate how many *grams of each product* would be produced by complete reaction of 12.5 g of the reactant indicated in boldface. Indicate clearly the mole ratio used for the conversion.
a. $TiBr_4(g) + \mathbf{H_2}(g) \rightarrow Ti(s) + HBr(g)$
b. $\mathbf{SiH_4}(g) + NH_3(g) \rightarrow Si_3N_4(s) + H_2(g)$
c. $NO(g) + \mathbf{H_2}(g) \rightarrow N_2(g) + 2H_2O(l)$
d. $\mathbf{Cu_2S}(s) \rightarrow Cu(s) + S(g)$

26. For each of the following *balanced* equations, calculate how many *grams of each product* would be produced by complete reaction of 15.0 g of the reactant indicated in boldface.
a. $\mathbf{2BCl_3}(s) + 3H_2(g) \rightarrow 2B(s) + 6HCl(g)$
b. $\mathbf{2Cu_2S}(s) + 3O_2(g) \rightarrow 2Cu_2O(s) + 2SO_2(g)$
c. $2Cu_2O(s) + \mathbf{Cu_2S}(s) \rightarrow 6Cu(s) + SO_2(g)$
d. $CaCO_3(s) + \mathbf{SiO_2}(s) \rightarrow CaSiO_3(s) + CO_2(g)$

27. Bottled propane is used in areas away from natural gas pipelines for cooking and heating, and is also the source of heat in most gas barbecue grills. Propane burns in oxygen according to the following balanced chemical equation:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

Calculate the mass in grams of water vapor produced if 3.11 mol of propane is burned.

28. Sulfuric acid is produced by first burning sulfur to produce sulfur trioxide gas

$$2S(s) + 3O_2(g) \rightarrow 2SO_3(g)$$

then dissolving the sulfur trioxide gas in water

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(l)$$

Calculate the mass of sulfuric acid produced if 1.25 g of sulfur is reacted as indicated in the above equations.

29. When elemental carbon is burned in the open atmosphere, with plenty of oxygen gas present, the product is carbon dioxide.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

However, when the amount of oxygen present during the burning of the carbon is restricted, carbon monoxide is more likely to result.

$$2C(s) + O_2(g) \rightarrow 2CO(g)$$

What mass of each product is expected when a 5.00-g sample of pure carbon is burned under each of these conditions?

30. If baking soda (sodium hydrogen carbonate) is heated strongly, the following reaction occurs:

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

Calculate the mass of sodium carbonate that will remain if a 1.52-g sample of sodium hydrogen carbonate is heated.

31. Although we usually think of substances as "burning" only in oxygen gas, the process of rapid oxidation to produce a flame may also take place in other strongly oxidizing gases. For example, when iron is heated and placed in pure chlorine gas, the iron "burns" according to the following (unbalanced) reaction:

$$Fe(s) + Cl_2(g) \rightarrow FeCl_3(s)$$

31. How many milligrams of iron(III) chloride result when 15.5 mg of iron is reacted with an excess of chlorine gas?

32. When yeast is added to a solution of glucose or fructose, the sugars are said to undergo *fermentation* and ethyl alcohol is produced.

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$

This is the reaction by which wines are produced from grape juice. Calculate the mass of ethyl alcohol, C_2H_5OH, produced when 5.25 g of glucose, $C_6H_{12}O_6$, undergoes this reaction.

33. Sulfurous acid is unstable in aqueous solution and gradually decomposes to water and sulfur dioxide gas (which explains the choking odor associated with sulfurous acid solutions).

$$H_2SO_3(aq) \rightarrow H_2O(l) + SO_2(g)$$

If 4.25 g of sulfurous acid undergoes this reaction, what mass of sulfur dioxide is released?

34. Small quantities of ammonia gas can be generated in the laboratory by heating an ammonium salt with a strong base. For example, ammonium chloride reacts with sodium hydroxide according to the following balanced equation:

$$NH_4Cl(s) + NaOH(s) \rightarrow NH_3(g) + NaCl(s) + H_2O(g)$$

What mass of ammonia gas is produced if 1.39 g of ammonium chloride reacts completely?

35. Elemental phosphorus burns in oxygen with an intensely hot flame, producing a brilliant light and clouds of the oxide product. These properties of the combustion of phosphorus have led to its being used in bombs and incendiary devices for warfare.

$$P_4(s) + 5O_2(g) \rightarrow 2P_2O_5(s)$$

If 4.95 g of phosphorus is burned, what mass of oxygen does it combine with?

36. Although we tend to make less use of mercury these days because of the environmental problems created by its improper disposal, mercury is still an important metal because of its unusual property of existing as a liquid at room temperature. One process by which mercury is produced industrially is through the heating of its common ore cinnabar (mercuric sulfide, HgS) with lime (calcium oxide, CaO).

$$4HgS(s) + 4CaO(s) \rightarrow 4Hg(l) + 3CaS(s) + CaSO_4(s)$$

What mass of mercury would be produced by complete reaction of 10.0 kg of HgS?

37. Ammonium nitrate has been used as a high explosive because it is unstable and decomposes into several gaseous substances. The rapid expansion of the gaseous substances produces the explosive force.

$$NH_4NO_3(s) \rightarrow N_2(g) + O_2(g) + H_2O(g)$$

Calculate the mass of each product gas if 1.25 g of ammonium nitrate reacts.

38. If common sugars are heated too strongly, they char as they decompose into carbon and water vapor. For example, if sucrose (table sugar) is heated, the reaction is

$$C_{12}H_{22}O_{11}(s) \rightarrow 12C(s) + 11H_2O(g)$$

What mass of carbon is produced if 1.19 g of sucrose decomposes completely?

39. Thionyl chloride, $SOCl_2$, is used as a very powerful drying agent in many synthetic chemistry experiments in which the presence of even small amounts of water would be detrimental. The unbalanced chemical equation is

$$SOCl_2(l) + H_2O(l) \rightarrow SO_2(g) + HCl(g)$$

Calculate the mass of water consumed by complete reaction of 35.0 g of $SOCl_2$.

40. Magnesium metal, which burns in oxygen with an intensely bright white flame, has been used in photographic flash units. The balanced equation for this reaction is

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

How many grams of MgO(s) are produced by complete reaction of 1.25 g of magnesium metal?

9.4 Calculations Involving a Limiting Reactant

QUESTIONS

41. Imagine you are chatting with a friend who has not yet taken a chemistry course. How would you explain the concept of *limiting reactant* to her? Your textbook uses the analogy of an automobile manufacturer ordering four wheels for each engine ordered as an example. Can you think of another analogy that might help your friend to understand the concept?

42. Explain how one determines which reactant in a process is the limiting reactant. Does this depend only on the masses of the reactant present? Is the mole ratio in which the reactants combine involved?

43. What is the *theoretical yield* for a reaction, and how does this quantity depend on the limiting reactant?

44. What does it mean to say a reactant is present "in excess" in a process? Can the *limiting reactant* be present in excess? Does the presence of an excess of a reactant affect the mass of products expected for a reaction?

PROBLEMS

45. For each of the following *unbalanced* reactions, suppose exactly 5.00 g of *each reactant* is taken. Determine which reactant is limiting, and also determine what mass of the excess reagent will remain after the limiting reactant is consumed.

a. $Na_2B_4O_7(s) + H_2SO_4(aq) + H_2O(l) \rightarrow H_3BO_3(s) + Na_2SO_4(aq)$

b. $CaC_2(s) + H_2O(l) \rightarrow Ca(OH)_2(s) + C_2H_2(g)$

c. $NaCl(s) + H_2SO_4(l) \rightarrow HCl(g) + Na_2SO_4(s)$
d. $SiO_2(s) + C(s) \rightarrow Si(l) + CO(g)$

46. For each of the following *unbalanced* chemical equations, suppose that exactly 5.00 g of *each* reactant is taken. Determine which reactant is limiting, and calculate what mass of each product is expected (assuming that the limiting reactant is completely consumed).

a. $S(s) + H_2SO_4(aq) \rightarrow SO_2(g) + H_2O(l)$
b. $MnO_2(s) + H_2SO_4(l) \rightarrow Mn(SO_4)_2(s) + H_2O(l)$
c. $H_2S(g) + O_2(g) \rightarrow SO_2(g) + H_2O(l)$
d. $AgNO_3(aq) + Al(s) \rightarrow Ag(s) + Al(NO_3)_3(aq)$

47. For each of the following *unbalanced* chemical equations, suppose 10.0 g of *each* reactant is taken. Show by calculation which reactant is the limiting reagent. Calculate the mass of each product that is expected.

a. $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
b. $Al(s) + Cl_2(g) \rightarrow AlCl_3(s)$
c. $NaOH(s) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$
d. $NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$

48. For each of the following *unbalanced* chemical equations, suppose that exactly 1.00 g of *each* reactant is taken. Determine which reactant is limiting, and calculate what mass of the product in boldface is expected (assuming that the limiting reactant is completely consumed).

a. $CS_2(l) + O_2(g) \rightarrow \mathbf{CO_2}(g) + SO_2(g)$
b. $NH_3(g) + CO_2(g) \rightarrow CN_2H_4O(s) + \mathbf{H_2O}(g)$
c. $H_2(g) + MnO_2(s) \rightarrow MnO(s) + \mathbf{H_2O}(g)$
d. $I_2(l) + Cl_2(g) \rightarrow \mathbf{ICl}(g)$

49. For each of the following *unbalanced* chemical equations, suppose 1.00 g of *each* reactant is taken. Show by calculation which reactant is limiting. Calculate the mass of each product that is expected.

a. $UO_2(s) + HF(aq) \rightarrow UF_4(aq) + H_2O(l)$
b. $NaNO_3(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + HNO_3(aq)$
c. $Zn(s) + HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
d. $B(OH)_3(s) + CH_3OH(l) \rightarrow B(OCH_3)_3(s) + H_2O(l)$

50. For each of the following *unbalanced* chemical equations, suppose 10.0 mg of *each* reactant is taken. Show by calculation which reactant is limiting. Calculate the mass of each product that is expected.

a. $CO(g) + H_2(g) \rightarrow CH_3OH(l)$
b. $Al(s) + I_2(s) \rightarrow AlI_3(s)$
c. $Ca(OH)_2(aq) + HBr(aq) \rightarrow CaBr_2(aq) + H_2O(l)$
d. $Cr(s) + H_3PO_4(aq) \rightarrow CrPO_4(s) + H_2(g)$

51. A demonstration many chemistry teachers like to perform in class is to combine aqueous lead nitrate solution with aqueous potassium iodide solution.

$$2KI(aq) + Pb(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + PbI(s)$$

Both reactants are colorless when the solutions are freshly prepared, but the solid product is bright yellow, so the process demonstrates a double displacement very effectively.

Suppose a solution containing 1.25 g of KI is combined with a solution containing 2.42 g of $Pb(NO_3)_2$. What mass of PbI would result?

52. An experiment that led to the formation of the new field of organic chemistry involved the synthesis of urea, CN_2H_4O, by the controlled reaction of ammonia and carbon dioxide:

$$2NH_3(g) + CO_2(g) \rightarrow CN_2H_4O(s) + H_2O(l)$$

What is the theoretical yield of urea when 100. g of ammonia is reacted with 100. g of carbon dioxide?

53. Lead(II) oxide from an ore can be reduced to elemental lead by heating in a furnace with carbon.

$$PbO(s) + C(s) \rightarrow Pb(l) + CO(g)$$

Calculate the expected yield of lead if 50.0 kg of lead oxide is heated with 50.0 kg of carbon.

54. If steel wool (iron) is heated until it glows and is placed in a bottle containing pure oxygen, the iron reacts spectacularly to produce iron(III) oxide.

$$Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$$

If 1.25 g of iron is heated and placed in a bottle containing 0.0204 mol of oxygen gas, what mass of iron(III) oxide is produced?

55. One method for chemical analysis involves finding some reagent that will precipitate the species of interest. The mass of the precipitate is then used to determine what mass of the species of interest was present in the original sample. For example, calcium ion can be precipitated from solution by addition of sodium oxalate. The balanced equation is

$$Ca^{2+}(aq) + Na_2C_2O_4(aq) \rightarrow CaC_2O_4(s) + 2Na^{+}(aq)$$

Suppose a solution is known to contain approximately 15 g of calcium ion. Show by calculation whether the addition of a solution containing 15 g of sodium oxalate will precipitate all of the calcium from the sample.

56. The copper(II) ion in a copper(II) sulfate solution reacts with potassium iodide to produce the triiodide ion, I_3^-. This reaction is commonly used to determine how much copper is present in a given sample.

$$CuSO_4(aq) + KI(aq) \rightarrow CuI(s) + KI_3(aq) + K_2SO_4(aq)$$

If 2.00 g of KI is added to a solution containing 0.525 g of $CuSO_4$, calculate the mass of each product produced.

57. Hydrogen peroxide is used as a cleaning agent in the treatment of cuts and abrasions for several reasons. It is an oxidizing agent that can directly kill many microorganisms; it decomposes upon contact with blood, releasing elemental oxygen gas (which inhibits the growth of anaerobic microorganisms); and it foams upon contact with blood, which provides a cleansing action. In the laboratory, small quantities

of hydrogen peroxide can be prepared by the action of an acid on an alkaline earth metal peroxide, such as barium peroxide.

$$BaO_2(s) + 2HCl(aq) \rightarrow H_2O_2(aq) + BaCl_2(aq)$$

What amount of hydrogen peroxide should result when 1.50 g of barium peroxide is treated with 25.0 mL of hydrochloric acid solution containing 0.0272 g of HCl per mL?

58. Silicon carbide, SiC, is one of the hardest materials known. Surpassed in hardness only by diamond, it is sometimes known commercially as carborundum. Silicon carbide is used primarily as an abrasive for sandpaper and is manufactured by heating common sand (silicon dioxide, SiO_2) with carbon in a furnace.

$$SiO_2(s) + C(s) \rightarrow CO(g) + SiC(s)$$

What mass of silicon carbide should result when 1.0 kg of pure sand is heated with an excess of carbon?

9.5 Percent Yield

QUESTIONS

59. Your text talks about several sorts of "yield" when experiments are performed in the laboratory. Students often confuse these terms. Define, compare, and contrast what are meant by *theoretical* yield, *actual* yield, and *percent* yield.

60. The text explains that one reason why the actual yield for a reaction may be less than the theoretical yield is side reactions. Suggest some other reasons why the percent yield for a reaction might not be 100%.

61. According to his prelaboratory theoretical yield calculations, a student's experiment should have produced 1.44 g of magnesium oxide. When he weighed his product after reaction, only 1.23 g of magnesium oxide was present. What is the student's percent yield?

62. Mercury used to be prepared in the laboratory by heating mercuric oxide.

$$2HgO(s) \rightarrow 2Hg(l) + O_2(g)$$

When 1.25 g of mercuric oxide is heated, what is the theoretical yield of mercury? Suppose 1.09 g of mercury is actually collected. What is the percent yield?

PROBLEMS

63. The compound sodium thiosulfate pentahydrate, $Na_2S_2O_3 \cdot 5H_2O$, is important commercially to the photography business as "hypo," because it has the ability to dissolve unreacted silver salts from photographic film during development. Sodium thiosulfate pentahydrate can be produced by boiling elemental sulfur in an aqueous solution of sodium sulfite.

$$S(s) + Na_2SO_3(aq) + H_2O(l) \rightarrow Na_2S_2O_3 \cdot 5H_2O(s)$$
(unbalanced)

What is the theoretical yield of sodium thiosulfate pentahydrate when 3.25 g of sulfur is boiled with 13.1 g of sodium sulfite? Sodium thiosulfate pentahydrate is very soluble in water. What is the percent yield of the synthesis if a student doing this experiment is able to isolate (collect) only 5.26 g of the product?

64. Alkali metal hydroxides are sometimes used to "scrub" excess carbon dioxide from the air in closed spaces (such as submarines and spacecraft). For example, lithium hydroxide reacts with carbon dioxide according to the unbalanced chemical equation

$$LiOH(s) + CO_2(g) \rightarrow Li_2CO_3(s) + H_2O(g)$$

Suppose a lithium hydroxide canister contains 155 g of LiOH(*s*). What mass of $CO_2(g)$ will the canister be able to absorb? If it is found that after 24 hours of use the canister has absorbed 102 g of carbon dioxide, what percentage of its capacity has been reached?

65. Although they were formerly called the inert gases, at least the heavier elements of Group 8 do form relatively stable compounds. For example, xenon combines directly with elemental fluorine at elevated temperatures in the presence of a nickel catalyst.

$$Xe(g) + 2F_2(g) \rightarrow XeF_4(s)$$

What is the theoretical mass of xenon tetrafluoride that should form when 130. g of xenon is reacted with 100. g of F_2? What is the percent yield if only 145 g of XeF_4 is actually isolated?

66. A common undergraduate laboratory analysis for the amount of sulfate ion in an unknown sample is to precipitate and weigh the sulfate ion as barium sulfate.

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

The precipitate produced, however, is very finely divided, and frequently some is lost during filtration before weighing. If a sample containing 1.12 g of sulfate ion is treated with 5.02 g of barium chloride, what is the theoretical yield of barium sulfate to be expected? If only 2.02 g of barium sulfate is actually collected, what is the percent yield?

Additional Problems

67. Natural waters often contain relatively high levels of calcium ion, Ca^{2+}, and hydrogen carbonate ion (bicarbonate), HCO_3^-, from the leaching of minerals into the water. When such water is used commercially or in the home, heating of the water leads to the formation of solid calcium carbonate, $CaCO_3$, which forms a deposit ("scale") on the interior of boilers, pipes, and other plumbing fixtures.

$$Ca(HCO_3)_2(aq) \rightarrow CaCO_3(s) + CO_2(g) + H_2O(l)$$

If a sample of well water contains 2.0×10^{-3} mg of $Ca(HCO_3)_2$ per milliliter, what mass of $CaCO_3$ scale would 1.0 mL of this water be capable of depositing?

68. One process for the commercial production of baking soda (sodium hydrogen carbonate) involves the following reaction, in which the carbon dioxide is used in its solid form ("dry ice") both to serve as a source of reactant and to cool the reaction system to a temperature low enough for the sodium hydrogen carbonate to precipitate:

$$NaCl(aq) + NH_3(aq) + H_2O(l) + CO_2(s) \rightarrow NH_4Cl(aq) + NaHCO_3(s)$$

Because they are relatively cheap, sodium chloride and water are typically present in excess. What is the expected yield of $NaHCO_3$ when one performs such a synthesis using 10.0 g of ammonia and 15.0 g of dry ice, with an excess of NaCl and water?

69. A favorite demonstration among chemistry instructors, to show that the properties of a compound differ from those of its constituent elements, involves iron filings and powdered sulfur. If the instructor takes samples of iron and sulfur and just mixes them together, the two elements can be separated from one another with a magnet (iron is attracted to a magnet, sulfur is not). If the instructor then combines and *heats* the mixture of iron and sulfur, a reaction takes place and the elements combine to form iron(II) sulfide (which is not attracted by a magnet).

$$Fe(s) + S(s) \rightarrow FeS(s)$$

Suppose 5.25 g of iron filings is combined with 12.7 g of sulfur. What is the theoretical yield of iron(II) sulfide?

70. When the sugar glucose, $C_6H_{12}O_6$, is burned in air, carbon dioxide and water vapor are produced. Write the balanced chemical equation for this process, and calculate the theoretical yield of carbon dioxide when 1.00 g of glucose is burned completely.

71. When elemental copper is strongly heated with sulfur, a mixture of CuS and Cu_2S is produced, with CuS predominating.

$$Cu(s) + S(s) \rightarrow CuS(s)$$
$$2Cu(s) + S(s) \rightarrow Cu_2S(s)$$

What is the theoretical yield of CuS when 31.8 g of Cu(*s*) is heated with 50.0 g of S? (Assume only CuS is produced in the reaction.) What is the percent yield of CuS if only 40.0 g of CuS can be isolated from the mixture?

72. Barium chloride solutions are used in chemical analysis for the quantitative precipitation of sulfate ion from solution.

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

Suppose a solution is known to contain on the order of 150 mg of sulfate ion. What mass of barium chloride should be added to guarantee precipitation of all the sulfate ion?

73. The traditional method of analysis for the amount of chloride ion present in a sample is to dissolve the sample in water and then slowly to add a solution of silver nitrate. Silver chloride is very insoluble in water, and by adding a slight excess of silver nitrate, it is possible effectively to remove all chloride ion from the sample.

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Suppose a 1.054-g sample is known to contain 10.3% chloride ion by mass. What mass of silver nitrate must be used to completely precipitate the chloride ion from the sample? What mass of silver chloride will be obtained?

74. For each of the following reactions, give the balanced equation for the reaction and state the meaning of the equation in terms of numbers of *individual molecules* and in terms of *moles* of molecules.

a. $UO_2(s) + HF(aq) \rightarrow UF_4(aq) + H_2O(l)$
b. $NaC_2H_3O_2(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + HC_2H_3O_2(aq)$
c. $Mg(s) + HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
d. $B_2O_3(s) + H_2O(l) \rightarrow B(OH)_3(aq)$

75. True or false? For the reaction represented by the balanced chemical equation

$$Mg(OH)_2(aq) + 2HCl(aq) \rightarrow 2H_2O(l) + MgCl_2(aq)$$

for 0.40 mol of $Mg(OH)_2$, 0.20 mol of HCl will be needed.

76. Consider the balanced equation

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

What mole ratio enables you to calculate the number of moles of oxygen needed to react exactly with a given number of moles of $C_3H_8(g)$? What mole ratios enable you to calculate how many moles of each product form from a given number of moles of C_3H_8?

77. For each of the following balanced reactions, calculate how many *moles of each product* would be produced by complete conversion of *0.50 mol* of the reactant indicated in boldface. Indicate clearly the mole ratio used for the conversion.

a. $\mathbf{2H_2O_2}(l) \rightarrow 2H_2O(l) + O_2(g)$
b. $\mathbf{2KClO_3}(s) \rightarrow 2KCl(s) + 3O_2(g)$
c. $\mathbf{2Al}(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
d. $\mathbf{C_3H_8}(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

78. For each of the following balanced equations, indicate how many *moles of the product* could be produced by complete reaction of *1.00 g* of the reactant indicated in boldface. Indicate clearly the mole ratio used for the conversion.

a. $\mathbf{NH_3}(g) + HCl(g) \rightarrow NH_4Cl(s)$
b. $\mathbf{CaO}(s) + CO_2(g) \rightarrow CaCO_3(s)$
c. $\mathbf{4Na}(s) + O_2(g) \rightarrow 2Na_2O(s)$
d. $\mathbf{2P}(s) + 3Cl_2(g) \rightarrow 2PCl_3(l)$

79. Using the average atomic masses given inside the front cover of the text, calculate how many *moles* of each substance the following masses represent.

a. 4.21 g of copper(II) sulfate
b. 7.94 g of barium nitrate
c. 1.24 mg of water
d. 9.79 g of tungsten
e. 1.45 lb of sulfur

ethyl alcohol, C_2H_5OH
of carbon

average atomic masses given inside the of the text, calculate the *mass in grams* following samples.

of nitric acid
mol of mercury
$\times 10^{-5}$ mol of potassium chromate
of aluminum chloride
mol of sulfur hexafluoride
of ammonia
mol of sodium peroxide

of the following *incomplete* and *unbalanced* indicate how many *moles* of the *second re-* be required to react completely with of the *first reactant.*

$+ H_2SO_4 \rightarrow$
$+ NaCl(aq) \rightarrow$
$(aq) + Na_2CO_3(aq) \rightarrow$
$+ O_2(g) \rightarrow$

in the commercial production of sulfuric, involves the conversion of sulfur diox- into sulfur trioxide, SO_3.

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

of SO_2 reacts completely, what mass of SO_3 result?

metals occur naturally as sulfide compounds; include ZnS and CoS. Air pollution often the processing of these ores, because dioxide is released as the ore is converted sulfide to the oxide by roasting (smelting). example, consider the unbalanced equation for reaction for zinc:

$$ZnS(s) + O_2(g) \rightarrow ZnO(s) + SO_2(g)$$

many kilograms of sulfur dioxide are produced 1.0×10^2 kg of ZnS is roasted in excess oxygen process?

peroxide is added to water, elemental oxy- is generated:

$$Na_2O_2(s) + H_2O(l) \rightarrow NaOH(aq) + O_2(g)$$

3.25 g of sodium peroxide is added to a large of water. What mass of oxygen gas will be produced?

elemental copper is placed in a solution of sil- nitrate, the following oxidation–reduction reac- takes place, forming elemental silver:

$$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

What mass of copper is required to remove all the from a silver nitrate solution containing 1.95 of silver nitrate?

When small quantities of elemental hydrogen gas are needed for laboratory work, the hydrogen is often generated by chemical reaction of a metal with acid. example, zinc reacts with hydrochloric acid, re- leasing gaseous elemental hydrogen:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

What mass of hydrogen gas is produced when 2.50 g of zinc is reacted with excess aqueous hydrochloric acid?

87. The gaseous hydrocarbon acetylene, C_2H_2, is used in welders' torches because of the large amount of heat released when acetylene burns with oxygen.

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(g)$$

How many grams of oxygen gas are needed for the complete combustion of 150 g of acetylene?

88. For each of the following *unbalanced* chemical equations, suppose exactly 5.0 g of each reactant is taken. Determine which reactant is limiting, and calculate what mass of each product is expected, assuming that the limiting reactant is completely consumed.

a. $Na(s) + Br_2(l) \rightarrow NaBr(s)$
b. $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$
c. $NH_4Cl(aq) + NaOH(aq) \rightarrow NH_3(g) + H_2O(l) + NaCl(aq)$
d. $Fe_2O_3(s) + CO(g) \rightarrow Fe(s) + CO_2(g)$

89. For each of the following *unbalanced* chemical equations, suppose 25.0 g of each reactant is taken. Show by calculation which reactant is limiting. Calculate the theoretical yield in grams of the product in boldface.

a. $C_2H_5OH(l) + O_2(g) \rightarrow \mathbf{CO_2}(g) + H_2O(l)$
b. $N_2(g) + O_2(g) \rightarrow \mathbf{NO}(g)$
c. $NaClO_2(aq) + Cl_2(g) \rightarrow ClO_2(g) + \mathbf{NaCl}(aq)$
d. $H_2(g) + N_2(g) \rightarrow \mathbf{NH_3}(g)$

90. Hydrazine, N_2H_4, emits a large quantity of energy when it reacts with oxygen, which has led to hydrazine's use as a fuel for rockets:

$$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(g)$$

How many moles of each of the gaseous products are produced when 20.0 g of pure hydrazine is ignited in the presence of 20.0 g of pure oxygen? How many grams of each product are produced?

91. Although elemental chlorine, Cl_2, is added to drinking water supplies primarily to kill microorganisms, another beneficial reaction that also takes place removes sulfides (which would impart unpleasant odors or tastes to the water). For example, the noxious-smelling gas hydrogen sulfide (its odor resembles that of rotten eggs) is removed from water by chlorine by the following reaction:

$$H_2S(aq) + Cl_2(aq) \rightarrow HCl(aq) + S_8(s) \quad \text{(unbalanced)}$$

What mass of sulfur is removed from the water when 50. L of water containing 1.5×10^{-5} g of H_2S per liter is treated with 1.0 g of $Cl_2(g)$?

92. Before going to lab, a student read in his lab manual that the percent yield for a difficult reaction to be studied was likely to be only 40.% of the theoretical yield. The student's prelab stoichiometric calculations predict that the theoretical yield should be 12.5 g. What is the student's actual yield likely to be?

CHAPTER 9

Chemical Quantities

CHAPTER ANSWERS

2. The coefficients of the balanced chemical equation for a reaction indicate the *relative numbers of moles* of each reactant that combine during the process as well as the number of moles of each product formed.

4. Balanced chemical equations tell us in what molar ratios substances combine to form products; not, in what mass proportions they combine.

6.

a. $(NH_4)_2CO_3(s) \rightarrow 2NH_3(g) + CO_2(g) + H_2O(g)$

One formula unit of solid ammonium carbonate decomposes to produce two molecules of ammonia gas, one molecule of carbon dioxide gas, and one molecule of water vapor. One mole of solid ammonium carbonate decomposes into two moles of gaseous ammonia, one mole of carbon dioxide gas, and one mole of water vapor.

b. $6Mg(s) + P_4(s) \rightarrow 2Mg_3P_2(s)$

Six atoms of magnesium metal react with one molecule of solid phosphorus (P_4) to make two formula units of solid magnesium phosphide. Six moles of magnesium metal react with one mole of phosphorus solid (P_4) to produce two moles of solid magnesium phosphide.

c. $4Si(s) + S_8(s) \rightarrow 2Si_2S_4(l)$

Four atoms of solid silicon react with one molecule of solid sulfur (S_8) to form two molecules of liquid disilicon tetrasulfide. Four moles of solid silicon react with one mole of solid sulfur (S_8) to form two moles of liquid disilicon tetrasulfide.

d. $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

One molecule of liquid ethanol burns with three molecules of oxygen gas to produce two molecules of carbon dioxide gas and three molecules of water vapor. One mole of liquid ethanol burns with three moles of oxygen gas to produce two moles of gaseous carbon dioxide and three moles of water vapor.

8. Balanced chemical equations tell us in what molar ratios substances combine to form products; not in what mass proportions they combine. How could 2 g of reactant produce a total of 3 g of products?

10. $S(s) + 2H_2SO_4(l) \rightarrow 3SO_2(g) + 2H_2O(l)$

For SO_2, $\left(\dfrac{3\ \text{mol}\ SO_2}{1\ \text{mol}\ S}\right)$

For H_2O, $\left(\dfrac{2\ \text{mol}\ H_2O}{1\ \text{mol}\ S}\right)$

For H_2SO_4, $\left(\dfrac{2\ \text{mol}\ H_2SO_4}{1\ \text{mol}\ S}\right)$

12.

a. $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$

$5.0\ \text{mol}\ C_2H_6 \times \dfrac{7\ \text{mol}\ O_2}{2\ \text{mol}\ C_2H_6} = 17.5\ \text{mol}\ O_2\ (18\ \text{mol}\ O_2)$

b. $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(g)$

$5.0\ \text{mol}\ P_4 \times \dfrac{5\ \text{mol}\ O_2}{1\ \text{mol}\ P_4} = 25\ \text{mol}\ O_2$

c. $CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$

$5.0\ \text{mol CaO} \times \dfrac{1\ \text{mol}\ CO_2}{1\ \text{mol CaO}} = 5.0\ \text{mol}\ CO_2$

d. $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

$5.0\ \text{mol Fe} \times \dfrac{3\ \text{mol}\ O_2}{4\ \text{mol Fe}} = 3.75\ \text{mol}\ O_2\ (3.8\ \text{mol}\ O_2)$

14.

a. $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

molar mass of NH_4Cl, 53.49 g

$0.50\ \text{mol}\ NH_3 \times \dfrac{1\ \text{mol}\ NH_4Cl}{1\ \text{mol}\ NH_3} = 0.50\ \text{mol}\ NH_4Cl$

$0.50\ \text{mol}\ NH_4Cl \times \dfrac{53.49\ \text{g}\ NH_4Cl}{1\ \text{mol}\ NH_4Cl} = 27\ \text{g}\ NH_4Cl$

b. $CH_4(g) + 4S(g) \rightarrow CS_2(l) + 2H_2S(g)$

molar masses: CS_2, 76.15 g; H_2S, 34.09 g

$$0.50 \text{ mol S} \times \frac{1 \text{ mol } CS_2}{4 \text{ mol S}} = 0.125 \text{ mol } CS_2 \ (= 0.13 \text{ mol } CS_2)$$

$$0.125 \text{ mol } CS_2 \times \frac{76.15 \text{ g } CS_2}{1 \text{ mol } CS_2} = 9.5 \text{ g } CS_2$$

$$0.50 \text{ mol S} \times \frac{2 \text{ mol } H_2S}{4 \text{ mol S}} = 0.25 \text{ mol } H_2S$$

$$0.25 \text{ mol } H_2S \times \frac{34.09 \text{ g } H_2S}{1 \text{ mol } H_2S} = 8.5 \text{ g } H_2S$$

c. $PCl_3(l) + 3H_2O(l) \rightarrow H_3PO_3(aq) + 3HCl(aq)$

molar masses: H_3PO_3, 81.99 g; HCl, 36.46 g

$$0.50 \text{ mol } PCl_3 \times \frac{1 \text{ mol H3}PO_3}{1 \text{ mol } PCl_3} = 0.50 \text{ mol } H_3PO_3$$

$$0.50 \text{ mol } H_3PO_3 \times \frac{81.99 \text{ g } H_3PO_3}{1 \text{ mol } H_3PO_3} = 41 \text{ g } H_3PO_3$$

$$0.50 \text{ mol } PCl_3 \times \frac{3 \text{ mol HCl}}{1 \text{ mol } PCl_3} = 1.5 \text{ mol HCl}$$

$$1.5 \text{ mol HCl} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 54.7 = 55 \text{ g HCl}$$

d. $NaOH(s) + CO_2(g) \rightarrow NaHCO_3(s)$

molar mass of $NaHCO_3$ = 84.01 g

$$0.50 \text{ mol NaOH} \times \frac{1 \text{ mol } NaHCO_3}{1 \text{ mol NaOH}} = 0.50 \text{ mol } NaHCO_3$$

$$0.50 \text{ mol } NaHCO_3 \times \frac{84.01 \text{ g } NaHCO_3}{1 \text{ mol } NaHCO_3} = 42 \text{ g } NaHCO_3$$

16. Before doing the calculations, the equations must be *balanced.*

a. $4KO_2(s) + 2H_2O(l) \rightarrow 3O_2(g) + 4KOH(s)$

$$0.625 \text{ mol KOH} \times \frac{3 \text{ mol } O_2}{4 \text{ mol KOH}} = 0.469 \text{ mol } O_2$$

b. $SeO_2(g) + 2H_2Se(g) \rightarrow 3Se(s) + 2H_2O(g)$

$$0.625 \text{ mol } H_2O \times \frac{3 \text{ mol Se}}{2 \text{ mol } H_2O} = 0.938 \text{ mol Se}$$

c. $2CH_3CH_2OH(l) + O_2(g) \rightarrow 2CH_3CHO(aq) + 2H_2O(l)$

$$0.625 \text{ mol } H_2O \times \frac{2 \text{ mol } CH_3CHO}{2 \text{ mol } H_2O} = 0.625 \text{ mol } CH_3CHO$$

d. $Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(l) + Al_2O_3(s)$

$$0.625 \text{ mol } Al_2O_3 \times \frac{2 \text{ mol Fe}}{1 \text{ mol } Al_2O_3} = 1.25 \text{ mol Fe}$$

[illegible]. Stoichiometry is the process of using a chemical equation to calculate the relative masses of reactants and products involved in a reaction.

[illegible].

a. molar mass $Li_2CO_3 = 73.89$ g; 2.36 mg = 0.00236 g

$$0.00236 \text{ g } Li_2CO_3 \times \frac{1 \text{ mol } Li_2CO_3}{73.89 \text{ g } Li_2CO_3} = 3.19 \times 10^{-5} \text{ mol } Li_2CO_3$$

b. molar mass U = 238.0 g

$$1.92 \times 10^{-3} \text{ g U} \times \frac{1 \text{ mol U}}{238.0 \text{ g U}} = 8.07 \times 10^{-6} \text{ mol U}$$

c. molar mass $PbCl_2 = 278.1$ g; 3.21 kg = 3.21×10^3 g

$$3.21 \times 10^3 \text{ g } PbCl_2 \times \frac{1 \text{ mol}}{278.1 \text{ g}} = 11.5 \text{ mol } PbCl_2$$

d. molar mass $C_6H_{12}O_6 = 180.2$ g

$$4.62 \text{ g } C_6H_{12}O_6 \times \frac{1 \text{ mol } C_6H_{12}O_6}{180.2 \text{ g } C_6H_{12}O_6} = 0.0256 \text{ mol } C_6H_{12}O_6$$

e. molar mass KOH = 56.11 g

$$7.75 \text{ g KOH} \times \frac{1 \text{ mol KOH}}{56.11 \text{ g KOH}} = 0.138 \text{ mol KOH}$$

22.

a. molar mass of CuI = 190.5 g

$$0.624 \text{ mol CuI} \times \frac{190.5 \text{ g CuI}}{1 \text{ mol CuI}} = 119 \text{ g CuI}$$

b. molar mass of $Br_2 = 159.8$ g

$$4.24 \text{ mol } Br_2 \times \frac{159.8 \text{ g } Br_2}{1 \text{ mol } Br_2} = 678 \text{ g } Br_2$$

c. molar mass of $XeF_4 = 207.4$ g

$$0.000211 \text{ mol } XeF_4 \times \frac{207.4 \text{ g } XeF_4}{1 \text{ mol } XeF_4} = 0.0438 \text{ g } XeF_4$$

d. molar mass of $C_2H_4 = 28.05$ g

$$9.11 \text{ mol } C_2H_4 \times \frac{28.05 \text{ g } C_2H_4}{1 \text{ mol } C_2H_4} = 256 \text{ g } C_2H_4$$

e. molar mass of $NH_3 = 17.03$ g; 1.21 millimol = 0.00121 mol

$$0.00121 \text{ mol } NH_3 \times \frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3} = 0.0206 \text{ g } NH_3$$

f. molar mass of NaOH = 40.00 g

$$4.25 \text{ mol NaOH} \times \frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}} = 170 \text{ g NaOH}$$

g. molar mass of KI = 166.0 g

$$1.27 \times 10^{-6} \text{ mol KI} \times \frac{166.0 \text{ g KI}}{1 \text{ mol KI}} = 2.11 \times 10^{-4} \text{ g KI}$$

24. Before any calculations are done, the equations must be *balanced.*

a. $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

molar mass $C_2H_5OH = 46.07$ g

$$5.00 \text{ g } C_2H_5OH \times \frac{1 \text{ mol } C_2H_5OH}{46.07 \text{ g } C_2H_5OH} \times \frac{3 \text{ mol } O_2}{1 \text{ mol } C_2H_5OH} = 0.326 \text{ mol } O_2$$

b. $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(g)$

molar mass $P_4 = 123.88$ g

$$5.00 \text{ g } P_4 \times \frac{1 \text{ mol } P_4}{123.88 \text{ g}} \times \frac{5 \text{ mol } O_2}{1 \text{ mol } P_4} = 0.202 \text{ mol } O_2$$

c. $MgO(s) + CO_2(g) \rightarrow MgCO_3(s)$

molar mass MgO = 40.31 g

$$5.00 \text{ g MgO} \times \frac{1 \text{ mol MgO}}{40.31 \text{ g MgO}} \times \frac{1 \text{ mol } CO_2}{1 \text{ mol MgO}} = 0.124 \text{ mol } CO_2$$

d. $2Fe(s) + O_2(g) \rightarrow 2FeO(s)$

molar mass Fe = 55.85 g

$$5.00 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{1 \text{ mol } O_2}{2 \text{ mol Fe}} = 0.0448 \text{ mol } O_2$$

a. $2BCl_3(s) + 3H_2(g) \rightarrow 2B(s) + 6HCl(g)$

molar masses: BCl_3, 117.16 g; B, 10.81 g; HCl, 36.46 g

$$15.0 \text{ g } BCl_3 \times \frac{1 \text{ mol } BCl_3}{117.16 \text{ g } BCl_3} = 0.128 \text{ mol } BCl_3$$

$$0.128 \text{ mol } BCl_3 \times \frac{2 \text{ mol B}}{2 \text{ mol } BCl_3} \times \frac{10.81 \text{ g B}}{1 \text{ mol B}} = 1.38 \text{ g B}$$

$$0.128 \text{ mol } BCl_3 \times \frac{6 \text{ mol HCl}}{2 \text{ mol } BCl_3} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 14.0 \text{ g HCl}$$

b. $2Cu_2S(s) + 3O_2(g) \rightarrow 2Cu_2O(s) + 2SO_2(g)$

molar masses: Cu_2S, 159.17 g; Cu_2O, 143.1 g; SO_2, 64.07 g

$$15.0 \text{ g } Cu_2S \times \frac{1 \text{ mol } Cu_2S}{159.17 \text{ g } Cu_2S} = 0.09424 \text{ mol } Cu_2S$$

$$0.09424 \text{ mol } Cu_2S \times \frac{2 \text{ mol } Cu_2O}{2 \text{ mol } Cu_2S} \times \frac{143.1 \text{ g } Cu_2O}{1 \text{ mol } Cu_2O} = 13.5 \text{ g } Cu_2O$$

$$0.09424 \text{ mol } Cu_2S \times \frac{2 \text{ mol } SO_2}{2 \text{ mol } Cu_2S} \times \frac{64.07 \text{ g } SO_2}{1 \text{ mol } SO_2} = 6.04 \text{ g } SO_2$$

c. $2Cu_2O(s) + Cu_2S(s) \rightarrow 6Cu(s) + SO_2(g)$

molar masses: Cu_2S, 159.17 g; Cu, 63.55 g; SO_2, 64.07 g

$$15.0 \text{ g } Cu_2S \times \frac{1 \text{ mol } Cu_2S}{159.17 \text{ g } Cu_2S} = 0.09424 \text{ mol } Cu_2S$$

$$0.09424 \text{ mol } Cu_2S \times \frac{6 \text{ mol Cu}}{1 \text{ mol } Cu_2S} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 35.9 \text{ g Cu}$$

$$0.09424 \text{ mol } Cu_2S \times \frac{1 \text{ mol } SO_2}{1 \text{ mol } Cu_2S} \times \frac{64.07 \text{ g } SO_2}{1 \text{ mol } SO_2} = 6.04 \text{ g } SO_2$$

d. $CaCO_3(s) + SiO_2(s) \rightarrow CaSiO_3(s) + CO_2(g)$

molar masses: SiO_2, 60.09 g; $CaSiO_3$, 116.17 g; CO_2, 44.01 g

$$15.0 \text{ g } SiO_2 \times \frac{1 \text{ mol } SiO_2}{60.09 \text{ g } SiO_2} = 0.2496 \text{ mol } SiO_2$$

$$0.2496 \text{ mol } SiO_2 \times \frac{1 \text{ mol } CaSiO_3}{1 \text{ mol } SiO_2} \times \frac{116.17 \text{ g } CaSiO_3}{1 \text{ mol } CaSiO_3} = 29.0 \text{ g } CaSiO_3$$

$$0.2496 \text{ mol } SiO_2 \times \frac{1 \text{ mol } CO_2}{1 \text{ mol } SiO_2} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 11.0 \text{ g } CO_2$$

28. $2S(s) + 3O_2(g) \rightarrow 2SO_3(g)$ $\quad$ $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(l)$

It would make things simpler if the first equation were expressed in terms of one mole of SO_3 since the second equation is expressed in terms of 1 mole of SO_3. To do this, divide the first equation by two:

$S(s) + \frac{3}{2} O_2(g) \rightarrow SO_3(g)$ $\quad$ $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(l)$

By doing this, we now have the simpler relationship that one mole of S will produce one mole of H_2SO_4.

molar masses: S, 32.07 g; H_2SO_4, 98.09 g

$$1.25 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 0.03898 \text{ mol S}$$

$$0.03988 \text{ mol S} \times \frac{1 \text{ mol } H_2SO_4}{1 \text{ mol S}} = 0.03988 \text{ mol } H_2SO_4$$

$$0.03988 \text{ mol } H_2SO_4 \times \frac{98.09 \text{ g } H_2SO_4}{1 \text{ mol } H_2SO_4} = 3.82 \text{ g } H_2SO_4$$

30. $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$

molar masses: $NaHCO_3$, 84.01 g; Na_2CO_3, 106.0 g

$$1.52 \text{ g } NaHCO_3 \times \frac{1 \text{ mol } NaHCO_3}{84.01 \text{ g } NaHCO_3} = 0.01809 \text{ mol } NaHCO_3$$

$$0.01809 \text{ mol } NaHCO_3 \times \frac{1 \text{ mol } Na_2CO_3}{2 \text{ mol } NaHCO_3} = 0.009047 \text{ mol } Na_2CO_3$$

$$0.009047 \text{ mol } Na_2CO_3 \times \frac{106.0 \text{ g } Na_2CO_3}{1 \text{ mol } Na_2CO_3} = 0.959 \text{ g } Na_2CO_3$$

32. $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$

molar masses: $C_6H_{12}O_6$, 180.2 g; C_2H_5OH, 46.07 g

$$5.25 \text{ g } C_6H_{12}O_6 \times \frac{1 \text{ mol } C_6H_{12}O_6}{180.2 \text{ g } C_6H_{12}O_6} = 0.02913 \text{ mol } C_6H_{12}O_6$$

$$0.02913 \text{ mol } C_6H_{12}O_6 \times \frac{2 \text{ mol } C_2H_5OH}{1 \text{ mol } C_6H_{12}O_6} = 0.5826 \text{ mol } C_2H_5OH$$

$$0.5286 \text{ mol } C_2H_5OH \times \frac{46.07 \text{ g } C_2H_5OH}{1 \text{ mol } C_2H_5OH} = 2.68 \text{ g ethyl alcohol}$$

34. $NH_4Cl(s) + NaOH(s) \rightarrow NH_3(g) + NaCl(s) + H_2O(g)$

molar masses: NH_4Cl, 53.49 g; NH_3, 17.03 g

$$1.39 \text{ g } NH_4Cl \times \frac{1 \text{ mol } NH_4Cl}{53.49 \text{ g } NH_4Cl} = 0.02599 \text{ mol } NH_4Cl$$

$$0.02599 \text{ mol } NH_4Cl \times \frac{1 \text{ mol } NH_3}{1 \text{ mol } NH_4Cl} = 0.02599 \text{ mol } NH_3$$

$$0.02599 \text{ mol } NH_3 \times \frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3} = 0.443 \text{ g } NH_3$$

36. $4HgS(s) + 4CaO(s) \rightarrow 4Hg(l) + 3CaS(s) + CaSO_4(s)$

molar masses: HgS, 232.7 g Hg, 200.6 g; 10.0 kg = 1.00×10^4 g

$$1.00 \times 10^4 \text{ g HgS} \times \frac{1 \text{ mol HgS}}{232.7 \text{ g HgS}} = 42.97 \text{ mol HgS}$$

$$42.97 \text{ mol HgS} \times \frac{4 \text{ mol Hg}}{4 \text{ mol HgS}} = 42.97 \text{ mol Hg}$$

$$42.97 \text{ mol Hg} \times \frac{200.6 \text{ g Hg}}{1 \text{ mol Hg}} = 8.62 \times 10^3 \text{ g Hg} = 8.62 \text{ kg Hg}$$

38. $C_{12}H_{22}O_{11}(s) \rightarrow 12C(s) + 11H_2O(g)$

molar masses: $C_{12}H_{22}O_{11}$, 342.3 g; C, 12.01

$$1.19 \text{ g } C_{12}H_{22}O_{11} \times \frac{1 \text{ mol } C_{12}H_{22}O_{11}}{342.3 \text{ g } C_{12}H_{22}O_{11}} = 3.476 \times 10^{-3} \text{ mol } C_{12}H_{22}O_{11}$$

$$3.476 \times 10^{-3} \text{ mol } C_{12}H_{22}O_{11} \times \frac{12 \text{ mol C}}{1 \text{ mol } C_{12}H_{22}O_{11}} = 0.04172 \text{ mol C}$$

$$0.04172 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.501 \text{ g C}$$

40. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

molar masses: Mg, 24.31 g; MgO, 40.31 g

$$1.25 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} = 5.14 \times 10^{-2} \text{ mol Mg}$$

$$5.14 \times 10^{-2} \text{ mol Mg} \times \frac{2 \text{ mol MgO}}{2 \text{ mol Mg}} = 5.14 \times 10^{-2} \text{ mol MgO}$$

$$5.14 \times 10^{-2} \text{ mol MgO} \times \frac{40.31 \text{ g MgO}}{1 \text{ mol MgO}} = 2.07 \text{ g MgO}$$

42. To determine the limiting reactant, first calculate the number of moles of each reactant present. Then determine how these numbers of moles correspond to the stoichiometric ratio indicated by the balanced chemical equation for the reaction.

44. A reactant is present *in excess* if there is more of that reactant present than is needed to combine with the limiting reactant for the process. By definition, the limiting reactant cannot be present in excess. An excess of any reactant does not affect the theoretical yield for a process; the theoretical yield is determined by the limiting reactant.

46.

a. $S(s) + 2H_2SO_4(aq) \rightarrow 3SO_2(g) + 2H_2O(l)$

Molar masses: S, 32.07 g; H_2SO_4, 98.09 g; SO_2, 64.07 g; H_2O, 18.02 g

$$5.00 \text{ g S} \times \frac{1 \text{ mol}}{32.07 \text{ g}} = 0.1559 \text{ mol S}$$

$$5.00 \text{ g } H_2SO_4 \times \frac{1 \text{ mol}}{98.09 \text{ g}} = 0.05097 \text{ mol } H_2SO_4$$

According to the balanced chemical equation, we would need twice as much sulfuric acid as sulfur for complete reaction of both reactants. We clearly have much less sulfuric acid present than sulfur; sulfuric acid is the limiting reactant. The calculation of the masses of products produced is based on the number of moles of the sulfuric acid.

$$0.05097 \text{ mol } H_2SO_4 \times \frac{3 \text{ mol } SO_2}{2 \text{ mol } H_2SO_4} \times \frac{64.07 \text{ g } SO_2}{1 \text{ mol } SO_2} = 4.90 \text{ g } SO_2$$

$$0.05097 \text{ mol } H_2SO_4 \times \frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2SO_4} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 0.918 \text{ g } H_2O$$

b. $MnO_2(s) + 2H_2SO_4(aq) \rightarrow Mn(SO_4)_2 + 2H_2O(l)$

molar masses: MnO_2, 86.94 g; H_2SO_4 98.09 g; $Mn(SO_4)_2$, 247.1 g; H_2O, 18.02 g

$$5.00 \text{ g } MnO_2 \times \frac{1 \text{ mol}}{86.94 \text{ g}} = 0.05751 \text{ mol } MnO_2$$

$$5.00 \text{ g } H_2SO_4 \times \frac{1 \text{ mol}}{98.09 \text{ g}} = 0.05097 \text{ mol } H_2SO_4$$

According to the balanced chemical equation, we would need twice as much sulfuric acid as manganese(IV) oxide for complete reaction of both reactants. We do not have this much sulfuric acid, so sulfuric acid must be the limiting reactant. The amount of each product produced will be based on the sulfuric acid reacting completely.

$$0.05097 \text{ mol } H_2SO_4 \times \frac{1 \text{ mol } Mn(SO_4)_2}{2 \text{ mol } H_2SO_4} \times \frac{247.1 \text{ g } Mn(SO_4)_2}{1 \text{ mol } Mn(SO_4)_2} = 6.30 \text{ g } Mn(SO_4)_2$$

$$0.05097 \text{ mol } H_2SO_4 \times \frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2SO_4} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 0.918 \text{ g } H_2O$$

c. $2H_2S(g) + 3O_2(g) \rightarrow 2SO_2(g) + 2H_2O(l)$

Molar masses: H_2S, 34.09 g; O_2, 32.00 g; SO_2, 64.07 g; H_2O, 18.02 g

$$5.00 \text{ g } H_2S \times \frac{1 \text{ mol}}{34.09 \text{ g}} = 0.1467 \text{ mol } H_2S$$

$$5.00 \text{ g } O_2 \times \frac{1 \text{ mol}}{32.00 \text{ g}} = 0.1563 \text{ mol } O_2$$

According to the balanced equation, we would need 1.5 times as much O_2 as H_2S for complete reaction of both reactants. We don't have that much O_2, so O_2 must be the limiting reactant that will control the masses of each product produced.

$$0.1563 \text{ mol } O_2 \times \frac{2 \text{ mol } SO_2}{3 \text{ mol } O_2} \times \frac{64.07 \text{ g } SO_2}{1 \text{ mol } SO_2} = 6.67 \text{ g } SO_2$$

$$0.1563 \text{ mol } O_2 \times \frac{2 \text{ mol } H_2O}{3 \text{ mol } O_2} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 1.88 \text{ g } H_2O$$

d. $3AgNO_3(aq) + Al(s) \rightarrow 3Ag(s) + Al(NO_3)_3(aq)$

Molar masses: $AgNO_3$, 169.9 g; Al, 26.98 g; Ag, 107.9 g; $Al(NO_3)_3$, 213.0 g

$$5.00 \text{ g } AgNO_3 \times \frac{1 \text{ mol}}{169.9 \text{ g}} = 0.02943 \text{ mol } AgNO_3$$

$$5.00 \text{ g Al} \times \frac{1 \text{ mol}}{26.98 \text{ g}} = 0.1853 \text{ mol Al}$$

According to the balanced chemical equation, we would need three moles of $AgNO_3$ for every mole of Al for complete reaction of both reactants. We in fact have fewer moles of $AgNO_3$ than aluminum, so $AgNO_3$ must be the limiting reactant. The amount of product produced is calculated from the number of moles of the limiting reactant present:

$$0.02943 \text{ mol } AgNO_3 \times \frac{3 \text{ mol Ag}}{3 \text{ mol } AgNO_3} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 3.18 \text{ g Ag}$$

$$0.02943 \text{ mol } AgNO_3 \times \frac{1 \text{ mol } Al(NO_3)_3}{3 \text{ mol } AgNO_3} \times \frac{213.0 \text{ g } Al(NO_3)_3}{1 \text{ mol } Al(NO_3)_3} = 2.09 \text{ g}$$

48.

a. $CS_2(l) + 3O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$

Molar masses: CS_2, 76.15 g; O_2, 32.00 g; CO_2, 44.01 g

$$1.00 \text{ g } CS_2 \times \frac{1 \text{ mol}}{76.15 \text{ g}} = 0.01313 \text{ mol } CS_2$$

$$1.00 \text{ g } O_2 \times \frac{1 \text{ mol}}{32.00 \text{ g}} = 0.03125 \text{ mol } O_2$$

From the balanced chemical equation, we would need three times as much oxygen as carbon disulfide for complete reaction of both reactants. We do not have this much oxygen, and so oxygen must be the limiting reactant.

$$0.03125 \text{ mol } O_2 \times \frac{1 \text{ mol } CO_2}{3 \text{ mol } O_2} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 0.458 \text{ g } CO_2$$

b. $2NH_3(g) + CO_2(g) \rightarrow CN_2H_4O(s) + H_2O(l)$

Molar masses: NH_3, 17.03 g; CO_2, 44.01 g; H_2O, 18.02 g

$$1.00 \text{ g } NH_3 \times \frac{1 \text{ mol}}{17.03 \text{ g}} = 0.05872 \text{ mol } NH_3$$

$$1.00 \text{ g } CO_2 \times \frac{1 \text{ mol}}{44.01 \text{ g}} = 0.02272 \text{ mol } CO_2$$

The balanced chemical equation tells us that we would need twice as many moles of ammonia as carbon dioxide for complete reaction of both reactants. We have *more* than this amount of ammonia present, so the reaction will be limited by the amount of carbon dioxide present.

$$0.02272 \text{ mol } CO_2 \times \frac{1 \text{ mol } H_2O}{1 \text{ mol } CO_2} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 0.409 \text{ g } H_2O$$

c. $H_2(g) + MnO_2(s) \rightarrow MnO(s) + H_2O(l)$

Molar masses: H_2, 2.016 g; MnO_2, 86.94 g; H_2O, 18.02 g

$$1.00 \text{ g } H_2 \times \frac{1 \text{ mol}}{2.016 \text{ g}} = 0.496 \text{ mol } H_2$$

$$1.00 \text{ g } MnO_2 \times \frac{1 \text{ mol}}{86.94 \text{ g}} = 0.0115 \text{ mol } MnO_2$$

Because the coefficients of both reactants in the balanced chemical equation are the same, we would need equal amounts of both reactants for complete reaction. Therefore manganese(IV) oxide must be the limiting reactant and controls the amount of product obtained.

$$0.0115 \text{ mol } MnO_2 \times \frac{1 \text{ mol } H_2O}{1 \text{ mol } MnO_2} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 0.207 \text{ g } H_2O$$

d. $I_2(s) + Cl_2(g) \rightarrow 2ICl(g)$

Molar masses: I_2, 253.8 g; Cl_2, 70.90 g; ICl, 162.35 g

$$1.00 \text{ g } I_2 \times \frac{1 \text{ mol}}{253.8 \text{ g}} = 0.00394 \text{ mol } I_2$$

$$1.00 \text{ g } Cl_2 \times \frac{1 \text{ mol}}{70.90 \text{ g}} = 0.0141 \text{ mol } Cl_2$$

From the balanced chemical equation, we would need equal amounts of I_2 and Cl_2 for complete reaction of both reactants. As we have much less iodine than chlorine, iodine must be the limiting reactant.

$$0.00394 \text{ mol } I_2 \times \frac{2 \text{ mol ICl}}{1 \text{ mol } I_2} \times \frac{162.35 \text{ g ICl}}{1 \text{ mol ICl}} = 1.28 \text{ g ICl}$$

50.

a. $CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$

CO is the limiting reactant; 11.4 mg CH_3OH

b. $2Al(s) + 3I_2(s) \rightarrow 2AlI_3(s)$

I_2 is the limiting reactant; 10.7 mg AlI_3

c. $Ca(OH)_2(aq) + 2HBr(aq) \rightarrow CaBr_2(aq) + 2H_2O(l)$

HBr is the limiting reactant; 12.4 mg $CaBr_2$; 2.23 mg H_2O

d. $2Cr(s) + 2H_3PO_4(aq) \rightarrow 2CrPO_4(s) + 3H_2(g)$

H_3PO_4 is the limiting reactant; 15.0 mg $CrPO_4$; 0.309 mg H_2

52. $2NH_3(g) + CO_2(g) \rightarrow CN_2H_4O(s) + H_2O(l)$

molar masses: NH_3, 17.03 g; CO_2, 44.01 g; CN_2H_4O, 60.06 g

$$100. \text{ g } NH_3 \times \frac{1 \text{ mol}}{17.03 \text{ g}} = 5.872 \text{ mol } NH_3$$

$$100. \text{ g } CO_2 \times \frac{1 \text{ mol}}{44.01 \text{ g}} = 2.272 \text{ mol } CO_2$$

CO_2 is the limiting reactant that determines the yield of product.

$$2.272 \text{ mol } CO_2 \times \frac{1 \text{ mol } CN_2H_4O}{1 \text{ mol } CO_2} \times \frac{60.06 \text{ g } CN_2H_4O}{1 \text{ mol } CN_2H_4O} = 136 \text{ g } CN_2H_4O$$

54. $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

Molar masses: Fe, 55.85 g; Fe_2O_3, 159.7 g

$$1.25 \text{ g Fe} \times \frac{1 \text{ mol}}{55.85 \text{ g}} = 0.0224 \text{ mol Fe present}$$

Calculate how many mol of O_2 are required to react with this amount of Fe

$$0.0224 \text{ mol Fe} \times \frac{3 \text{ mol } O_2}{4 \text{ mol Fe}} = 0.0168 \text{ mol } O_2$$

Because we have more O_2 than this, Fe must be the limiting reactant.

$$0.0224 \text{ mol Fe} \times \frac{2 \text{ mol } Fe_2O_3}{4 \text{ mol Fe}} \times \frac{159.7 \text{ g } Fe_2O_3}{1 \text{ mol } Fe_2O_3} = 1.79 \text{ g } Fe_2O_3$$

56. $2CuSO_4(aq) + 5KI(aq) \rightarrow 2CuI(s) + KI_3(aq) + 2K_2SO_4(aq)$

molar masses: $CuSO_4$, 159.6 g; KI, 166.0 g; CuI, 190.5 g; KI_3, 419.8 g; K_2SO_4, 174.3 g

$$0.525 \text{ g } CuSO_4 \times \frac{1 \text{ mol}}{159.6 \text{ g}} = 3.29 \times 10^{-3} \text{ mol } CuSO_4$$

$$2.00 \text{ g KI} \times \frac{1 \text{ mol}}{166.0 \text{ g}} = 0.0120 \text{ mol KI}$$

To determine the limiting reactant, let's calculate what amount of KI would be needed to react with the given amount of $CuSO_4$ present.

$$3.29 \times 10^{-3} \text{ mol } CuSO_4 \times \frac{5 \text{ mol KI}}{2 \text{ mol } CuSO_4} = 8.23 \times 10^{-3} \text{ mol KI}$$

As we have more KI present than the amount required to react with the $CuSO_4$ present, $CuSO_4$ must be the limiting reactant that will control the amount of products produced.

$$3.29 \times 10^{-3} \text{ mol } CuSO_4 \times \frac{2 \text{ mol CuI}}{2 \text{ mol } CuSO_4} \times \frac{190.5 \text{ g CuI}}{1 \text{ mol CuI}} = 0.627 \text{ g CuI}$$

$$3.29 \times 10^{-3} \text{ mol } CuSO_4 \times \frac{1 \text{ mol } KI_3}{2 \text{ mol } CuSO_4} \times \frac{419.8 \text{ g } KI_3}{1 \text{ mol } KI_3} = 0.691 \text{ g } KI_3$$

$$3.29 \times 10^{-3} \text{ mol } CuSO_4 \times \frac{2 \text{ mol } K_2SO_4}{2 \text{ mol } CuSO_4} \times \frac{174.3 \text{ g } K_2SO_4}{1 \text{ mol } K_2SO_4} = 0.573 \text{ g } K_2SO_4$$

58. $SiO_2(s) + 3C(s) \rightarrow 2CO(g) + SiC(s)$

molar masses: SiO_2, 60.09 g; SiC, 40.10 g; 1.0 kg = 1.0×10^3 g

$$1.0 \times 10^3 \text{ g } SiO_2 \times \frac{1 \text{ mol}}{60.09 \text{ g}} = 16.64 \text{ mol } SiO_2$$

From the balanced chemical equation, if 16.64 mol of SiO_2 were to react completely (an excess of carbon is present), then 16.64 mol of SiC should be produced (the coefficients of SiO_2 and SiC are the same).

$$16.64 \text{ mol SiC} \times \frac{40.01 \text{ g}}{1 \text{ mol}} = 6.7 \times 10^2 \text{ g SiC} = 0.67 \text{ kg SiC}$$

60. If the reaction is performed in a solvent, the product may have substantial solubility in the solvent and the reaction may come to equilibrium before the full yield of product is achieved (See Chapter 17.). Loss of product may occur through operator error.

62. $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

molar masses: HgO, 216.6 g; Hg, 200.6 g

$$1.25 \text{ g HgO} \times \frac{1 \text{ mol}}{216.6 \text{ g}} = 0.005771 \text{ mol HgO}$$

$$0.005771 \text{ mol HgO} \times \frac{2 \text{ mol Hg}}{2 \text{ mol HgO}} \times \frac{200.6 \text{ g Hg}}{1 \text{ mol Hg}} = 1.16 \text{ g (theoretical yield)}$$

$$\% \text{ yield} = \frac{1.09 \text{ g actual}}{1.16 \text{ g theoretical}} \times 100 = 94.0\% \text{ of theory}$$

64. $2LiOH(s) + CO_2(g) \rightarrow Li_2CO_3(s) + H_2O(g)$

molar masses: LiOH, 23.95 g; CO_2, 44.01 g

$$155 \text{ g LiOH} \times \frac{1 \text{ mol LiOH}}{23.95 \text{ g LiOH}} \times \frac{1 \text{ mol } CO_2}{2 \text{ mol LiOH}} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 142 \text{ g } CO_2$$

As the cartridge has only absorbed 102 g CO_2 out of a total capacity of 142 g CO_2, the cartridge has absorbed

$$\frac{102 \text{ g}}{142 \text{ g}} \times 100 = 71.8\% \text{ of its capacity.}$$

66. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

molar masses: SO_4^{2-}, 96.07 g; $BaCl_2$, 208.2 g; $BaSO_4$, 233.4 g

$$1.12 \text{ g } SO_4^{2-} \times \frac{1 \text{ mol}}{96.07 \text{ g}} = 0.01166 \text{ mol } SO_4^{2-}$$

$$5.02 \text{ g } BaCl_2 \times \frac{1 \text{ mol}}{208.2 \text{ g}} = 0.02411 \text{ mol } BaCl_2 = 0.02411 \text{ mol } Ba^{2+}$$

SO_4^{2-} is the limiting reactant.

$$0.01166 \text{ mol } SO_4^{2-} \times \frac{1 \text{ mol } BaSO_4}{1 \text{ mol } SO_4^{2-}} \times \frac{233.4 \text{ g } BaSO_4}{1 \text{ mol } BaSO_4} = 2.72 \text{ g } BaSO_4$$

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{2.02 \text{ g}}{2.72 \text{ g}} \times 100 = 74.3\%$$

68. $NaCl(aq) + NH_3(aq) + H_2O(l) + CO_2(s) \rightarrow NH_4Cl(aq) + NaHCO_3(s)$

molar masses: NH_3, 17.03 g; CO_2, 44.01 g; $NaHCO_3$, 84.01 g

$$10.0 \text{ g } NH_3 \times \frac{1 \text{ mol}}{17.03 \text{ g}} = 0.5872 \text{ mol } NH_3$$

$$15.0 \text{ g } CO_2 \times \frac{1 \text{ mol}}{44.01 \text{ g}} = 0.3408 \text{ mol } CO_2$$

CO_2 is the limiting reactant.

$$0.3408 \text{ mol } CO_2 \times \frac{1 \text{ mol } NaHCO_3}{1 \text{ mol } CO_2} = 0.3408 \text{ mol } NaHCO_3$$

$$0.3408 \text{ mol } NaHCO_3 \times \frac{84.01 \text{ g}}{1 \text{ mol}} = 28.6 \text{ g } NaHCO_3$$

70. $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$

molar masses: glucose, 180.2 g; CO_2, 44.01 g

$1.00 \text{ g glucose} \times = 5.549 \times 10^{-3} \text{ mol glucose}$

$$5.549 \times 10^{-3} \text{ mol glucose} \times \frac{6 \text{ mol } CO_2}{1 \text{ mol glucose}} = 3.33 \times 10^{-2} \text{ mol } CO_2$$

$$3.33 \times 10^{-2} \text{ mol } CO_2 \times \frac{44.01 \text{ g}}{1 \text{ mol}} = 1.47 \text{ g } CO_2$$

72. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

millimolar ionic masses: Ba^{2+}, 137.3 mg; SO_4^{2-}, 96.07 mg; $BaCl_2$, 208.2 mg

$$150 \text{ mg } SO_4^{2-} \times \frac{1 \text{ mmol}}{96.07 \text{ mg}} = 1.56 \text{ millimol } SO_4^{2-}$$

As barium ion and sulfate ion react on a 1:1 stoichiometric basis, then 1.56 millimol of barium ion is needed, which corresponds to 1.56 millimol of $BaCl_2$.

$$1.56 \text{ millimol } BaCl_2 \times \frac{208.2 \text{ mg}}{1 \text{ mmol}} = 325 \text{ milligrams } BaCl_2 \text{ needed}$$

74.

a. $UO_2(s) + 4HF(aq) \rightarrow UF_4(aq) + 2H_2O(l)$

One molecule (formula unit) of uranium(IV) oxide will combine with four molecules of hydrofluoric acid, producing one uranium(IV) fluoride molecule and two water molecules. One mole of uranium(IV) oxide will combine with four moles of hydrofluoric acid to produce one mole of uranium(IV) fluoride and two moles of water.

b. $2NaC_2H_3O_2(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2HC_2H_3O_2(aq)$

Two molecules (formula units) of sodium acetate react exactly with one molecule of sulfuric acid, producing one molecule (formula unit) of sodium sulfate and two molecules of acetic acid. Two moles of sodium acetate will combine with one mole of sulfuric acid, producing one mole of sodium sulfate and two moles of acetic acid.

c. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

One magnesium atom will react with two hydrochloric acid molecules (formula units) to produce one molecule (formula unit) of magnesium chloride and one molecule of hydrogen gas. One mole of magnesium will combine with two moles of hydrochloric acid, producing one mole of magnesium chloride and one mole of gaseous hydrogen.

d. $B_2O_3(s) + 3H_2O(l) \rightarrow 2B(OH)_3(aq)$

One molecule of diboron trioxide will react exactly with three molecules of water, producing two molecules of boron trihydroxide (boric acid). One mole of diboron trioxide will combine with three moles of water to produce two moles of boron trihydroxide (boric acid).

76. For O_2: $\left(\frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8}\right)$ For CO_2: $\left(\frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8}\right)$ For H_2O: $\left(\frac{4 \text{ mol } H_2O}{1 \text{ mol } C_3H_8}\right)$

78.

a. $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

molar mass of NH_3 = 17.01 g

$$1.00 \text{ g } NH_3 \times \frac{1 \text{ mol}}{17.01 \text{ g}} = 0.0588 \text{ mol } NH_3$$

$$0.0588 \text{ mol } NH_3 \times \frac{1 \text{ mol } NH_4Cl}{1 \text{ mol } NH_3} = 0.0588 \text{ mol } NH_4Cl$$

b. $CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$

molar mass CaO = 56.08 g

$$1.00 \text{ g CaO} \times \frac{1 \text{ mol}}{56.08 \text{ g}} = 0.0178 \text{ mol CaO}$$

$$0.0178 \text{ mol CaO} \times \frac{1 \text{ mol } CaCO_3}{1 \text{ mol CaO}} = 0.0178 \text{ mol } CaCO_3$$

c. $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

molar mass Na = 22.99 g

$$1.00 \text{ g Na} \times \frac{1 \text{ mol}}{22.99 \text{ g}} = 0.0435 \text{ mol Na}$$

$$0.0435 \text{ mol Na} \times \frac{2 \text{ mol } Na_2O}{4 \text{ mol Na}} = 0.0217 \text{ mol } Na_2O$$

d. $2P(s) + 3Cl_2(g) \rightarrow 2PCl_3(l)$

molar mass P = 30.97 g

$$1.00 \text{ g P} \times \frac{1 \text{ mol}}{30.97 \text{ g}} = 0.0323 \text{ mol P}$$

$$0.0323 \text{ mol P} \times \frac{2 \text{ mol PCl}_3}{2 \text{ mol P}} = 0.0323 \text{ mol PCl}_3$$

80.

a. molar mass HNO_3 = 63.0 g

$$5.0 \text{ mol HNO}_3 \times \frac{63.0 \text{ g}}{1 \text{ mol}} = 3.2 \times 10^2 \text{ g HNO}_3$$

b. molar mass Hg = 200.6 g

$$0.000305 \text{ mol Hg} \times \frac{200.6 \text{ g}}{1 \text{ mol}} = 0.0612 \text{ g Hg}$$

c. molar mass K_2CrO_4 = 194.2 g

$$2.31 \times 10^{-5} \text{ mol K}_2\text{CrO}_4 \times \frac{194.2 \text{ g}}{1 \text{ mol}} = 4.49 \times 10^{-3} \text{ g K}_2\text{CrO}_4$$

d. molar mass $AlCl_3$ = 133.3 g

$$10.5 \text{ mol AlCl}_3 \times \frac{133.3 \text{ g}}{1 \text{ mol}} = 1.40 \times 10^3 \text{ g AlCl}_3$$

e. molar mass SF_6 = 146.1 g

$$4.9 \times 10^4 \text{ mol SF}_6 \times \frac{146.1 \text{ g}}{1 \text{ mol}} = 7.2 \times 10^6 \text{ g SF}_6$$

f. molar mass NH_3 = 17.01 g

$$125 \text{ mol NH}_3 \times \frac{17.01 \text{ g}}{1 \text{ mol}} = 2.13 \times 10^3 \text{ g NH}_3$$

g. molar mass Na_2O_2 = 77.98 g

$$0.01205 \text{ mol Na}_2\text{O}_2 \times \frac{77.98 \text{ g}}{1 \text{ mol}} = 0.9397 \text{ g Na}_2\text{O}_2$$

82. $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

molar masses: SO_2, 64.07 g; SO_3, 80.07 g; 150 kg = 1.5×10^5 g

$$1.5 \times 10^5 \text{ g } SO_2 \times \frac{1 \text{ mol}}{64.07 \text{ g}} = 2.34 \times 10^3 \text{ mol } SO_2$$

$$2.34 \times 10^3 \text{ mol } SO_2 \times \frac{2 \text{ mol } SO_3}{2 \text{ mol } SO_2} = 2.34 \times 10^3 \text{ mol } SO_3$$

$$2.34 \times 10^3 \text{ mol } SO_3 \times \frac{80.07 \text{ g}}{1 \text{ mol}} = 1.9 \times 10^5 \text{ g } SO_3 = 1.9 \times 10^2 \text{ kg } SO_3$$

84. $2Na_2O_2(s) + 2H_2O(l) \rightarrow 4NaOH(aq) + O_2(g)$

molar masses: Na_2O_2, 77.98 g; O_2, 32.00 g

$$3.25 \text{ g } Na_2O_2 \times \frac{1 \text{ mol}}{77.98 \text{ g}} = 0.0417 \text{ mol } Na_2O_2$$

$$0.0417 \text{ mol } Na_2O_2 \times \frac{1 \text{ mol } O_2}{2 \text{ mol } Na_2O_2} = 0.0209 \text{ mol } O_2$$

$$0.0209 \text{ mol } O_2 \times \frac{32.00 \text{ g}}{1 \text{ mol}} = 0.669 \text{ g } O_2$$

86. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

molar masses: Zn, 65.38 g; H_2, 2.016 g

$$2.50 \text{ g Zn} \times \frac{1 \text{ mol}}{65.38 \text{ g}} = 0.03824 \text{ mol Zn}$$

$$0.03824 \text{ mol Zn} \times \frac{1 \text{ mol } H_2}{1 \text{ mol Zn}} = 0.03824 \text{ mol } H_2$$

$$0.03824 \text{ mol } H_2 \times \frac{2.016 \text{ g}}{1 \text{ mol}} = 0.0771 \text{ g } H_2$$

88.

a. $2Na(s) + Br_2(l) \rightarrow 2NaBr(s)$

molar masses: Na, 22.99 g; Br_2, 159.8 g; NaBr, 102.9 g

$$5.0 \text{ g Na} \times \frac{1 \text{ mol}}{22.99 \text{ g}} = 0.2175 \text{ mol Na}$$

$$5.0 \text{ g } Br_2 \times \frac{1 \text{ mol}}{159.8 \text{ g}} = 0.03129 \text{ mol } Br_2$$

Intuitively, we would suspect that Br_2 is the limiting reactant because there is much less Br_2 than Na on a mole basis. To *prove* that Br_2 is the limiting reactant, the following calculation is needed:

$$0.03129 \text{ mol } Br_2 \times \frac{2 \text{ mol Na}}{1 \text{ mol } Br_2} = 0.06258 \text{ mol Na.}$$

Clearly there is more Na than this present, so Br_2 limits the reaction extent and the amount of NaBr formed.

$$0.03129 \text{ mol } Br_2 \times \frac{2 \text{ mol NaBr}}{1 \text{ mol } Br_2} = 0.06258 \text{ mol NaBr}$$

$$0.06258 \text{ mol NaBr} \times \frac{102.9 \text{ g}}{1 \text{ mol}} = 6.4 \text{ g NaBr}$$

b. $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

molar masses: Zn, 65.38 g; Cu, 63.55 g; $ZnSO_4$, 161.5 g; $CuSO_4$, 159.6 g

$$5.0 \text{ g Zn} \times \frac{1 \text{ mol}}{65.38 \text{ g}} = 0.07648 \text{ mol Zn}$$

$$5.0 \text{ g } CuSO_4 \times \frac{1 \text{ mol}}{159.6 \text{ g}} = 0.03132 \text{ mol } CuSO_4$$

As the coefficients of Zn and $CuSO_4$ are the *same* in the balanced chemical equation, an equal number of moles of Zn and $CuSO_4$ would be needed for complete reaction. There is less $CuSO_4$ present, so $CuSO_4$ must be the limiting reactant.

$$0.03132 \text{ mol } CuSO_4 \times \frac{1 \text{ mol } ZnSO_4}{1 \text{ mol } CuSO_4} = 0.03132 \text{ mol } ZnSO_4$$

$$0.03132 \text{ mol } ZnSO_4 \times \frac{161.5 \text{ g}}{1 \text{ mol}} = 5.1 \text{ g } ZnSO_4$$

$$0.03132 \text{ mol } CuSO_4 \times \frac{1 \text{ mol Cu}}{1 \text{ mol } CuSO_4} = 0.03132 \text{ mol Cu}$$

$$0.03132 \text{ mol Cu} \times \frac{63.55 \text{ g}}{1 \text{ mol}} = 2.0 \text{ g Cu}$$

c. $NH_4Cl(aq) + NaOH(aq) \rightarrow NH_3(g) + H_2O(l) + NaCl(aq)$

molar masses: NH_4Cl, 53.49 g; NaOH, 40.00 g; NH_3, 17.03 g; H_2O, 18.02 g; NaCl, 58.44 g

$$5.0 \text{ g } NH_4Cl \times \frac{1 \text{ mol}}{53.49 \text{ g}} = 0.09348 \text{ mol } NH_4Cl$$

$$5.0 \text{ g NaOH} \times \frac{1 \text{ mol}}{40.00 \text{ g}} = 0.1250 \text{ mol NaOH}$$

As the coefficients of NH_4Cl and NaOH are both *one* in the balanced chemical equation for the reaction, an equal number of moles of NH_4Cl and NaOH would be needed for complete reaction. There is less NH_4Cl present, so NH_4Cl must be the limiting reactant.

As the coefficients of the products in the balanced chemical equation are also all *one*, if 0.09348 mol of NH_4Cl (the limiting reactant) reacts completely, then 0.09348 mol of each product will be formed.

$$0.09348 \text{ mol } NH_3 \times \frac{17.03 \text{ g}}{1 \text{ mol}} = 1.6 \text{ g } NH_3$$

$$0.09348 \text{ mol } H_2O \times \frac{18.02 \text{ g}}{1 \text{ mol}} = 1.7 \text{ g } H_2O$$

$$0.09348 \text{ mol NaCl} \times \frac{58.44 \text{ g}}{1 \text{ mol}} = 5.5 \text{ g NaCl}$$

d. $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

molar masses: Fe_2O_3, 159.7 g; CO, 28.01 g; Fe, 55.85 g; CO_2, 44.01 g

$$5.0 \text{ g } Fe_2O_3 \times \frac{1 \text{ mol}}{159.7 \text{ g}} = 0.03131 \text{ mol } Fe_2O_3$$

$$5.0 \text{ g CO} \times \frac{1 \text{ mol}}{28.01 \text{ g}} = 0.1785 \text{ mol CO}$$

Because there is considerably less Fe_2O_3 than CO on a mole basis, let's see if Fe_2O_3 is the limiting reactant.

$$0.03131 \text{ mol } Fe_2O_3 \times \frac{3 \text{ mol CO}}{1 \text{ mol } Fe_2O_3} = 0.09393 \text{ mol CO}$$

There is 0.1785 mol of CO present, but we have determined that only 0.09393 mol CO would be needed to react with all the Fe_2O_3 present, so Fe_2O_3 must be the limiting reactant. CO is present in excess.

$$0.03131 \text{ mol } Fe_2O_3 \times \frac{2 \text{ mol Fe}}{1 \text{ mol } Fe_2O_3} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 3.5 \text{ g Fe}$$

$$0.03131 \text{ mol } Fe_2O_3 \times \frac{3 \text{ mol } CO_2}{1 \text{ mol } Fe_2O_3} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 4.1 \text{ g } CO_2$$

90. $N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(g)$

molar masses: N_2H_4, 32.05 g; O_2, 32.00 g; N_2, 28.02 g; H_2O, 18.02 g

$$20.0 \text{ g } N_2H_4 \times \frac{1 \text{ mol}}{32.05 \text{ g}} = 0.624 \text{ mol } N_2H_4$$

$$20.0 \text{ g } O_2 \times \frac{1 \text{ mol}}{32.00 \text{ g}} = 0.625 \text{ mol } O_2$$

The two reactants are present in nearly the required ratio for complete reaction (due to the 1:1 stoichiometry of the reaction and the very similar molar masses of the substances). We will consider N_2H_4 as the limiting reactant in the following calculations.

$$0.624 \text{ mol } N_2H_4 \times \frac{1 \text{ mol } N_2}{1 \text{ mol } N_2H_4} \times \frac{28.02 \text{ g } N_2}{1 \text{ mol } N_2} = 17.5 \text{ g } N_2$$

$$0.624 \text{ mol } N_2H_4 \times \frac{2 \text{ mol } H_2O}{1 \text{ mol } N_2H_4} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 22.5 \text{ g } H_2O$$

92. $12.5 \text{ g theory} \times \frac{40 \text{ g actual}}{100 \text{ g theory}} = 5.0 \text{ g}$