CHAPTER 9

Chemical Quantities

INTRODUCTION

In this chapter you will perform many chemical calculations, all of which are based on fundamental principles such as balanced equations. A balanced equation can provide more information than is apparent at first glance. You can use a balanced equation to help answer such questions as "How much is produced?" and "How much would be needed to make this amount?" Only a balanced equation will provide correct answers to these questions.

Often, when two reactants are mixed together, one of them will run out before the other one is all used up. In a situation like this, the amount of product you can make will be limited by the reactant that is used up first. The balanced equation will help you determine which reactant runs out and how much product you can make.

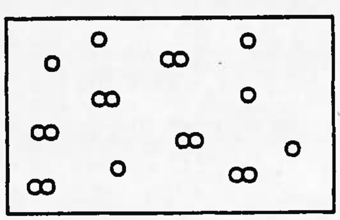
CHAPTER DISCUSSION

A main point of this chapter is to be able to calculate the mass of reactant needed or mass of product formed given the mass of one or more reactants or products (this is called stoichiometry). Realize that there is really nothing new in this chapter, but you are expected to put together what you have learned in the last few chapters. For example, you need to know how to balance an equation and what it means; you need to understand the mole concept; you need to be able to calculate molar masses of molecules. If you are having difficulty with any of these, make sure to get help (and re-read these chapters) or you will find this chapter quite difficult.

Also, make sure not to get lost in the math with stoichiometry problems. By now you should be used to thinking about these problems at a molecular level. Make sure you understand what you are solving for and how you are getting there. Look back to the percent by mass problem at the end of the Chapter Discussion for Chapter Eight in this *Study Guide*. Remember that dimensional analysis is not a help if you don't understand the problem (a shortcut is not a help if you get lost along the way).

In the text, "maps" are frequently used to show how to solve these problems. Don't merely memorize these, but instead make sure to understand the thinking behind them. Note that the balanced equation is always a part of any of these maps; make sure to understand why.

One way to make sense of stoichiometry is to first consider a molecular-level sketch of a reaction. For example, consider the following problem, and answer it before you read on.

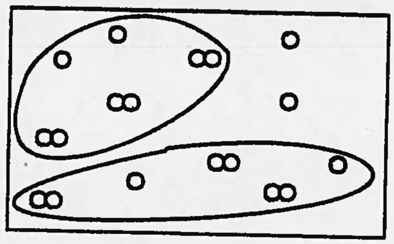
> The equation for a reaction is $2S + 3O_2 \rightarrow 2SO_3$. Consider the mixture of S and O_2 in a closed container as illustrated below:
>
> 
>
> ← This represents the entire container.
>
> Sketch a molecular-level representation of the product mixture.

To answer this question we need to take note of two things:

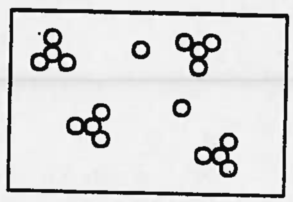
1. The number of molecules of each reactant given.
2. The ratio of the reactants that are needed.

The first piece of information is given in the problem; that is, there are six molecules of each reactant given. The second piece of information comes from the balanced equation; that is, for every two molecules of sulfur (S), three molecules of oxygen (O_2) are needed. Another way of stating this is for every two moles of sulfur (S), three moles of oxygen (O_2) are needed.

We can visualize this ratio by circling the reactants that react with each other as shown below:



The reactants produce the compound sulfur trioxide (SO_3), and we can represent the product mixture (including any leftover reactant) as



We can see that four molecules of SO_3 are produced, and two atoms of S are left over (unreacted). Make sure to understand this problem because it covers the basic concepts of stoichiometry. It even considers limiting reactants, which many students find to be the most difficult section of this chapter. The rest of stoichiometry has to do with the math (mass-mole conversions, essentially), and we will consider this briefly later.

While sketching these pictures is a good way of thinking about the problems initially, it is rather inefficient to solve all problems this way, and you will eventually want to formalize the solution a bit more. Before reading on, though, make sure to understand this example.

Formalizing a Solution

Let's consider the same problem, but different (and perhaps more efficient) ways of solving it. Recall the problem:

> You react 6 mol of S with 6 mol of O_2 according to the equation
>
> $$2S + 3O_2 \rightarrow 2SO_3$$
>
> Calculate the number of moles of SO_3 produced and the number of moles of leftover reactant.

Solution I

One way to solve this problem is to determine the number of moles of product formed if each reactant reacted completely. That is, change the given to two other problems:

1. How many moles of SO_3 could be produced from 6 mol of S and excess O_2?
2. How many moles of SO_3 could be produced from 6 mol of O_2 and excess S?

To answer this question we still need to know (as with the molecular-level sketch solution earlier) the number of moles we have (given in the problem) and the ratio from the balanced equation. Answer the two questions above before reading on.

From 6 mol of S, we can produce 6 mol of SO_3. There are a few ways to solve this with ratios, dimensional analysis, or even by inspection (the mole ratio between S and SO_3 is 2:2 or 1:1, thus for every 6 mol of S reacted, 6 mol of SO_3 will be produced).

From 6 mol of O_2, we can produce 4 mol of SO_3. Again, solve this using ratios or dimensional analysis. For example:

$$6 \text{ mol } O_2 \times \left(\frac{2 \text{ mol } SO_3}{3 \text{ mol } O_2}\right) = 4 \text{ mol } SO_3$$

So now we have two answers. That is, we have calculated 6 mol SO_3 and 4 mol SO_3. We know from our molecular level sketch that the answer is 4 mol (Note that it is NOT 10 mol—we do NOT simply add up the answers. Why not?). Let's make sense of this.

Realize what these two answers represent. They are the maximum amount of product that could be produced if the given reactant is used up completely. The answer must be the smaller number (4 mol in this case) because there is not enough O_2 to form 6 mol of SO_3. Once 4 mol of SO_3 are produced, there is no O_2 left, thus no more SO_3 can be produced. We now know, then, that 4 mol of SO_3 can be produced, and that O_2 is the limiting reactant; that is, the reactant that limits the reaction or which that runs out first.

So if all of the O_2 is used up, how much of the S is used? How much of the S is left over? We can answer these questions similarly to the previous question. We know that all 6 mol of O_2 are reacted. So how many moles of sulfur (S) would this require? Recall that the ratio between S and O_2 is 2:3 (from the balanced equation). Thus, we get

$$6 \text{ mol } O_2 \times \left(\frac{2 \text{ mol } SO_3}{3 \text{ mol } O_2}\right) = 4 \text{ mol S}$$

What does the 4 mol sulfur (S) represent? It represents the number of moles of S that are required to react with the 6 mol of O_2 that we know reacts. Thus we initially had 6 mol of sulfur, and 4 mol of S were reacted. So how many moles are left? Two moles of sulfur, just as we saw with our molecular-level sketches. This is one way of formalizing this problem–calculate the moles of product that would form if each reactant went to completion, and decide which reactant is limiting.

Solution II

There is another way we can solve this problem, and that is by comparing what we are given to what is needed for a complete reaction. For example, recall the problem

> You react 6 mol of S with 6 mol of O_2 according to the equation
>
> $$2S + 3O_2 \rightarrow 2SO_3$$
>
> Calculate the number of moles of SO_3 produced and the number of moles of leftover reactant.

We are given that we HAVE 6 mol of S and 6 mol of O_2. Can we determine which reactant is limiting without solving for the product twice?

The way to do this is to calculate the moles of each reactant that is needed to react with the moles of the other reactant that we are given. For example, we know that we have 6 mol of sulfur (S). How many moles of oxygen would be required to react completely with these 6 mol? Try this before reading on.

You should be able to calculate that 9 mol of oxygen (O_2) are needed to react with 6 mol of sulfur (S). Use the mole ratio given in the balanced equation to do so. If you are still having difficulty doing this, you need to talk with your instructor.

We know from the last solution (Solution I) that 4 mol of S are required to react with 6 mol of O_2. We can present this information in the following table:

	Moles Sulfur (S)	**Moles Oxygen (O_2)**
Have	6 mol S	6 mol O_2
Need	4 mol S	9 mol O_2

Note that we have more moles of sulfur than we need and fewer moles of O_2 than we need. Thus, O_2 must be the limiting reactant (the reactant that runs out first). We will therefore use the oxygen data to calculate the moles of product (SO_3) formed. That is

$$6 \text{ mol } O_2 \times \left(\frac{2 \text{ mol } SO_3}{3 \text{ mol } O_2}\right) = 4 \text{ moles } SO_3$$

Also, we can see from the table above that we have 6 mol of sulfur and need 4 mol of sulfur, thus 2 mol of sulfur are left over. This also agrees with our previous solution. It also agrees with our next solution, as we shall see.

Solution III

Another way of thinking about this problem is to set up a table that includes all the information shown in Solutions I and II. For example, consider the following

	2S	+	$3O_2$	→	$2SO_3$
Initial	6		6		0
Change	– ?		– ?		+ ?
End	?		?		?

Note that the 6 mol of each reactant (and no product initially) are represented in the "Initial" row. The "Change" row represents how much of each chemical reacts or is produced. The "End" row represents what constitutes the final reaction mixture.

Because we are assuming that the reaction goes to completion, we know that one (or possibly both) of the values for sulfur or oxygen must be zero (0) in the end row; that is, we "run out" of one (or possibly both) of the reactants. But which one(s)? We can decide this by realizing a crucial point:

The change row ratio has to be the same as the ratio of the coefficients in the balanced equation.

Make sure to understand this. The balanced equation represents the ratio of the reactants that react and the products that are formed, and the change row represents the same thing. Let's look, then, at the two possibilities:

	2S	+	$3O_2$	→	$2SO_3$
Initial	6		6		0
Change	–6		–9		+6
End	0		–3		6

	$2S$	+	$3O_2$	$\rightarrow$	$2SO_3$
Initial	6		6		0
Change	–4		–6		+4
End	2		0		4

In the first table, we are assuming that all the sulfur is reacted, and in the second example we are assuming that all the oxygen is used. We have already looked at other solutions to this problem, so we know that the answers to the second table are correct. But we can see why sulfur is not limiting by looking at the first table. For all of the sulfur to react, we need three more moles of oxygen than we have. We cannot end up with a negative amount of oxygen, so the first table must be incorrect.

Using a table such as this is convenient because all of the possible information is conveyed; we now know which reactant is limiting, how much of the excess reactant is leftover, and how much product is formed. It also emphasizes an understanding of what a balanced equation means because we have to use the ratio for the balanced equation in the "change" row.

Stoichiometry Problems with Masses

Once you understand the concepts of stoichiometry, the rest is math. Most typical problems will give you mass data of reactants, for example, and ask for the mass of products formed. As an example, consider the following problem:

> Hydrogen gas (H_2) reacts with oxygen gas (O_2) to form water (H_2O). If you react 10.0 g of hydrogen gas with 10.0 g of oxygen gas, what mass of water can be produced? How much of which reactant is left over?

Try to solve this problem before reading on. Remember, think about the problem before merely plugging numbers into an equation.

One way to start is to determine the balanced equation for the reaction. We know that we will need to know the mole ratio of the reactants and products to solve this problem.

You should get: $2H_2 + O_2 \rightarrow 2H_2O$ as the balanced equation.

This equation tells us that for every 2 mol of hydrogen gas we need 1 mol of oxygen gas to make 2 mol of water. We need to know how many moles of each reactant we have. How do we do this? By now you should know how to convert from grams to moles. In this case, you should be able to calculate the following values:

moles H_2: 4.96 moles

moles O_2: 0.313 moles

If you are having difficulty getting these numbers, review molar mass in Chapter 8 of your text or talk with an instructor.

Now that we know the number of moles of each reactant, we can solve the questions that were asked in the problem. You should be able to calculate the following:

mass of water formed: 11.3 g

mass of hydrogen left over: 8.74 g

If you are having difficulty with this, review the various solutions to the previous problem in this *Study Guide*.

LEARNING REVIEW

1. Rewrite the equation below in terms of moles of reactants and products.

   6.022×10^{23} molecules $H_2(g)$ + 6.022×10^{23} molecules $I_2(g) \rightarrow 1.204 \times 10^{24}$ molecules $HI(g)$

2. How many moles of hydrogen gas could be produced from 0.8 mol sodium and an excess of water? Solve this problem by writing the equation using moles and by using the mole ratio for sodium and hydrogen.

   $$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

3. How many moles of aluminum oxide could be produced from 0.12 mol Al?

   $$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

4. How many moles of zinc chloride would be formed from the reaction of 1.38 mol Zn with HCl?

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

5. Solid silver carbonate decomposes to produce silver metal, oxygen gas and carbon dioxide.

   a. Write a balanced chemical equation for this reaction.

   b. What mass of silver will be produced by the decomposition of 6.32 g silver carbonate?

6. When aqueous solutions of sodium sulfate and lead(II) nitrate are mixed, a solid white precipitate is formed. How much solid lead(II) sulfate could be produced from 12.0 g Na_2SO_4 if $Pb(NO_3)_2$ is in excess?

   $$Na_2SO_4(aq) + Pb(NO_3)_2(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$$

7. Hydrogen gas and chlorine gas will combine to produce gaseous hydrogen chloride. How many molecules of hydrogen chloride can be produced from 20.1 g hydrogen gas and excess chlorine gas?

8. Some lightweight backpacking stoves use kerosene as a fuel. Kerosene is composed of carbon and hydrogen, and although it is a mixture of molecules, we can represent the formula of kerosene as $C_{11}H_{24}$. When a kerosene stove is lit, the fuel reacts with oxygen in the air to produce carbon dioxide gas and water vapor. If it takes 15 g of kerosene to fry a trout for dinner, how many grams of water are produced?

   $$C_{11}H_{24}(l) + 17O_2(g) \rightarrow 11CO_2(g) + 12H_2O(g)$$

9. You are trying to prepare 6 copies of a three-page report. If you have on hand 6 copies of pages one and two, and 4 copies of page three

   a. How many complete reports can you produce?

   b. Which page limits the number of complete reports you can produce?

10. Manganese(IV) oxide reacts with hydrochloric acid to produce chlorine gas, manganese(II) chloride and water.

    $$MnO_2(s) + 4HCl(aq) \rightarrow Cl_2(g) + MnCl_2(aq) + 2H_2O(l)$$

    a. When 10.2 g MnO_2 react with 18.3 g HCl, which is the limiting reactant?

    b. What mass of chlorine gas can be produced?

    c. How many molecules of water can be produced?

11. The acid–base reaction between phosphoric acid and magnesium hydroxide produces solid magnesium phosphate and liquid water. If 121.0 g of phosphoric acid reacts with 89.70 g magnesium hydroxide, how many grams of magnesium phosphate will be produced?

12. If 85.6 g of potassium iodide reacts with 2.41×10^{24} molecules of chlorine gas, how many grams of iodine can be produced?

$$Cl_2(g) + 2KI(s) \rightarrow 2KCl(s) + I_2(s)$$

13. Aqueous sodium iodide reacts with aqueous lead(II) nitrate to produce the yellow precipitate lead(II) iodide and aqueous sodium nitrate.

    a. What is the theoretical yield of lead iodide if 125.5 g of sodium iodide reacts with 205.6 g of lead nitrate?
    164.5

    b. If the actual yield from this reaction is [illegible] g lead iodide, what is the percent yield?

ANSWERS TO LEARNING REVIEW

1. 6.022×10^{23} molecules is equivalent to 1 mol of molecules, and 1.204×10^{24} molecules is equivalent to $2(6.022 \times 10^{23}$ molecules), so the equation can be rewritten as

$$1 \text{ mol } H_2(g) + 1 \text{ mol } I_2(g) \rightarrow 2 \text{ mol } HI(g)$$

2. The balanced equation tells us that two moles of sodium react with two moles of water to form two moles of sodium hydroxide and four moles of hydrogen. By using mole ratios determined from the balanced equation, we can calculate the number of moles of reactants required and products produced from 0.8 mol sodium.

$$0.8 \text{ mol Na} \times \frac{2 \text{ mol } H_2O}{2 \text{ mol Na}} = 0.8 \text{ mol } H_2O$$ 0.8 mol sodium requires 0.8 mol H_2O.

$$0.8 \text{ mol Na} \times \frac{2 \text{ mol NaOH}}{2 \text{ mol Na}} = 0.8 \text{ mol NaOH}$$ 0.8 mol Na produces 0.8 mol NaOH.

$$0.8 \text{ mol Na} \times \frac{1 \text{ mol } H_2}{2 \text{ mol Na}} = 0.4 \text{ mol } H_2$$ 0.8 mol Na produces 0.4 mol H_2.

We can write the molar values we have calculated in equation form.

$$0.8 \text{ mol Na}(s) + 0.8 \text{ mol } H_2O(l) \rightarrow 0.8 \text{ mol NaOH}(aq) + 0.4 \text{ mol } H_2(g)$$

3. First make sure the equation is balanced. You should always determine whether or not an equation is balanced, and balance it if necessary. To solve this problem, we need to know the mole ratio for aluminum and aluminum oxide. The mole ratio represents the relationship between the mol of substance given in the problem and the mol of the desired substance and is taken directly from the balanced equation. The mole ratio for aluminum oxide and aluminum is.

$$\frac{2 \text{ mol } Al_2O_3}{4 \text{ mol Al}}$$

$$0.12 \text{ mol Al} \times \frac{2 \text{ mol } Al_2O_3}{4 \text{ mol Al}} = 0.060 \text{ mol } Al_2O_3$$

4. First make sure the equation is balanced. The mole ratio for zinc and zinc chloride is taken from the balanced equation and is.

$$\frac{1 \text{ mol } ZnCl_2}{1 \text{ mol Zn}}$$

$$1.38 \text{ mol Zn} \times \frac{1 \text{ mol } ZnCl_2}{1 \text{ mol Zn}} = 1.38 \text{ mol } ZnCl_2$$

Because there is a 1:1 mole ratio of $ZnCl_2$ to Zn, the number of moles of zinc equals the moles of zinc chloride produced.

5.

a. First write the formulas for reactants and products. Include the physical states. Then balance the equation.

$$2Ag_2CO_3(s) \rightarrow 4Ag(s) + O_2(g) + 2CO_2(g)$$

b. It is not possible to solve this problem by converting directly from grams of Ag_2CO_3 to grams of Ag. However, the balanced equation tells us the relationship between Ag_2CO_3 and Ag in moles. If we can convert grams of Ag_2CO_3 to moles, we can use the mole ratio to tell us how many moles of Ag are produced. To convert grams of Ag_2CO_3 to moles, you can produce a conversion factor from the equivalence statement that relates the number of moles to molar mass. The correct conversion factor is.

$$\frac{1 \text{ mol } Ag_2CO_3}{275.75 \text{ g } Ag_2CO_3}$$

$$6.32 \text{ g } Ag_2CO_3 \times \frac{1 \text{ mol } Ag_2CO_3}{275.75 \text{ g } Ag_2CO_3} = 0.0229 \text{ mol } Ag_2CO_3$$

Now we can use the mole ratio for Ag_2CO_3 and Ag to calculate the moles of Ag.

$$0.0229 \text{ mol } Ag_2CO_3 \times \frac{4 \text{ mol Ag}}{2 \text{ mol } Ag_2CO_3} = 0.0458 \text{ mol Ag}$$

We now know the moles of Ag, but we want to know the grams of Ag. The conversion factor below, which is derived from the molar mass of silver, will allow us to calculate grams.

$$\frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}}$$

$$0.0458 \text{ mol Ag} \times \frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}} = 4.94 \text{ g Ag}$$

If we string together all the parts of this problem, we can see that the overall strategy is to convert grams to moles using the molar mass, then moles to moles using the mole ratio, and moles to mass using the molar mass.

$$\underset{\substack{\uparrow \\ \text{grams of} \\ \text{reactant}}}{6.32 \text{ g } Ag_2CO_3} \times \underset{\substack{\uparrow \\ \text{molar mass} \\ \text{of reactant}}}{\frac{1 \text{ mol } Ag_2CO_3}{175.75 \text{ g } Ag_2CO_3}} \times \underset{\substack{\uparrow \\ \text{mole ratio}}}{\frac{4 \text{ mol Ag}}{2 \text{ mol } Ag_2CO_3}} \times \underset{\substack{\uparrow \\ \text{molar mass} \\ \text{of product}}}{\frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}}} = \underset{\substack{\uparrow \\ \text{grams of} \\ \text{product}}}{4.94 \text{ g Ag}}$$

6. This question provides us with grams of reactant and asks for grams of product. Because we are told that $Pb(NO_3)_2$ is in excess, the limiting reactant must be Na_2SO_4. The amount of precipitate that can be formed is determined by the amount of Na_2SO_4 relative to the grams of $PbSO_4$. We must first calculate the moles of Na_2SO_4, then use the mole ratio derived from the balanced equation to tell us how many moles of $PbSO_4$ are produced, and finally, we can use the molar mass of $PbSO_4$ to calculate the grams of $PbSO_4$.

$$12.0 \text{ g } Na_2SO_4 \times \frac{1 \text{ mol } Na_2SO_4}{142.05 \text{ g } Na_2SO_4} \times \frac{1 \text{ mol } PbSO_4}{1 \text{ mol } Na_2SO_4} \times \frac{303.27 \text{ g } PbSO_4}{1 \text{ mol } PbSO_4} = 25.6 \text{ g } PbSO_4$$

7. First write the balanced equation for this reaction.

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

This problem gives us grams of hydrogen and asks for molecules of hydrogen chloride. There is no way to convert grams of hydrogen directly to molecules of hydrogen chloride. However, we can convert grams of hydrogen to moles of hydrogen using the molar mass of hydrogen gas. The balanced equation provides a mole ratio so that we can calculate the moles of hydrogen chloride. Converting from moles to molecules can be done because we know that 1 mol of hydrogen chloride equals 6.022×10^{23} molecules of hydrogen chloride.

$$20.1 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \times \frac{2 \text{ mol HCl}}{1 \text{ mol } H_2} \times \frac{6.022 \times 10^{23} \text{ molecules HCl}}{1 \text{ mol HCl}}$$
$$= 1.20 \times 10^{25} \text{ molecules}$$

8. We are given grams of kerosene and asked for grams of water vapor. Because we cannot convert directly between grams of kerosene and grams of water, we first convert grams of kerosene to moles of kerosene using the molar mass of kerosene. Then use the mole ratio of kerosene and water from the balanced equation to determine the moles of water vapor. The molar mass of water will allow us to convert moles to grams of water.

$$15 \text{ g } C_{11}H_{24} \times \frac{1 \text{ mol } C_{11}H_{24}}{156.30 \text{ g } C_{11}H_{24}} \times \frac{12 \text{ mol } H_2O}{1 \text{ mol } C_{11}H_{24}} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 21 \text{ g } H_2O$$

9.

a. You can prepare 4 complete copies. Copies 5 and 6 would lack page three.

b. Page three limits the number of complete reports that can be produced.

10.

a. By looking at the grams of MnO_2 and the grams of HCl, it is impossible to tell which is the limiting reactant. It is possible to compare moles of reactants because we know the mole ratio of reactants from the balanced equation. So calculate the number of moles of each reactant. Then determine how many moles of product could be produced from each of the two reactants. The reactant that allows the fewest number of moles of product is the limiting reactant.

$$10.2 \text{ g } MnO_2 \times \frac{1 \text{ mol } MnO_2}{86.94 \text{ g } MnO_2} \times \frac{1 \text{ mol } Cl_2}{1 \text{ mol } MnO_2} = 0.117 \text{ mol } Cl_2$$

$$18.3 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \frac{1 \text{ mol } Cl_2}{4 \text{ mol HCl}} = 0.125 \text{ mol } Cl_2$$

From 10.2 MnO_2, 0.117 mol Cl_2 can be produced, and from 18.3 g HCl, 0.125 mol Cl_2 can be produced. So the limiting reactant is MnO_2.

b. We already know that the most chlorine we can make is 0.117 mol. To convert from moles to grams, use the molar mass of a chlorine molecule.

$$0.117 \text{ mol } Cl_2 \times \frac{70.90 \text{ g } Cl_2}{1 \text{ mol } Cl_2} = 8.30 \text{ g } Cl_2$$

c. The limiting reactant is manganese(IV) oxide, so we need to calculate the moles of water that can be produced from 10.2 g MnO_2. By using the mole ratio from the balanced equation we can calculate the moles of water. To convert from moles of water to the number of molecules, use Avogadro's number as a conversion factor.

$$10.2 \text{ g } MnO_2 \times \frac{1 \text{ mol } MnO_2}{86.94 \text{ g } MnO_2} \times \frac{2 \text{ mol } H_2O}{1 \text{ mol } MnO_2} \times \frac{6.022 \times 10^{23} \text{ molecules } H_2O}{1 \text{ mol } H_2O}$$
$$= 1.41 \times 10^{23} \text{ molecules } H_2O$$

11. First write the balanced equation for this reaction.

$$2H_3PO_4(aq) + 3Mg(OH)_2(s) \rightarrow Mg_3(PO_4)_2(s) + 6H_2O(l)$$

When the mass is given for two reactants,, and you are asked to determine the quantity of product that can be produced, you must first determine which reactant is limiting. Determine how many moles of product would be produced from each reactant. The reactant that will produce the fewest number of moles of product is the limiting reactant.

$$121.0 \text{ g } H_3PO_4 \times \frac{1 \text{ mol } H_3PO_4}{97.99 \text{ g } H_3PO_4} \times \frac{1 \text{ mol } Mg_3(PO_4)_2}{2 \text{ mol } H_3PO_4} = 0.6174 \text{ mol } Mg_3(PO_4)_2$$

$$89.70 \text{ g } Mg(OH)_2 \times \frac{1 \text{ mol } Mg(OH)_2}{58.33 \text{ g } Mg(OH)_2} \times \frac{1 \text{ mol } Mg_3(PO_4)_2}{3 \text{ mol } Mg(OH)_2} = 0.5126 \text{ mol } Mg_3(PO_4)_2$$

In this reaction, the $Mg(OH)_2$ is the limiting reactant. We now know how many moles of $Mg_3(PO_4)_2$ are produced, but we want to know the number of grams. Use the molar mass of $Mg(PO_4)_2$ to convert from moles of grams.

$$0.5126 \text{ mol } Mg_3(PO4)_2 \times \frac{262.87 \text{ g } Mg_3(PO4)_2}{1 \text{ mol } Mg_3(PO4)_2} = 134.7 \text{ g } Mg_3(PO4)_2$$

12. In this problem we are given quantities of two reactants, one expressed in grams and the other in molecules. Before we can calculate grams of product, we need to know which reactant limits the amount of product that can be produced. Convert the grams of KI to moles using the molar mass of KI, and calculate the moles of I_2 from the mole ratio.

$$85.6 \text{ g KI} \times \frac{1 \text{ mol KI}}{166.0 \text{ g KI}} \times \frac{1 \text{ mol } I_2}{2 \text{ mol KI}} = 0.258 \text{ mol } I_2$$

The quantity of the other reactant, Cl_2, is given in molecules, not grams. We can convert molecules of Cl_2 to moles of Cl_2 using Avogadro's number, 1 mol Cl_2 = 6.022×10^{23} molecules Cl_2.

$$2.41 \times 10^{24} \text{ molecules } Cl_2 \times \frac{1 \text{ mol } Cl_2}{6.022 \times 10^{23} \text{ molecules } Cl_2} = 4.00 \text{ mol } Cl_2$$

So 2.41×10^{24} molecules is equivalent to 4.00 moles of Cl_2. Now we can find the moles of I_2 that can be produced from 4.00 moles of Cl_2.

$$4.00 \text{ mol } Cl_2 \times \frac{1 \text{ mol } I_2}{1 \text{ mol } Cl_2} = 4.00 \text{ mol } I_2$$

KI limits the amount of I_2 that can be produced, so it is the limiting reactant. We can calculate the grams of I_2 using the molar mass of I_2.

$$0.258 \text{ mol } I_2 \times \frac{253.8 \text{ g } I_2}{1 \text{ mol } I_2} = 65.5 \text{ g } I_2$$

13.

a. First balance the equation.

$$2\ NaI + Pb(NO_3)_2 \rightarrow PbI_2 + 2\ NaNO_3$$

This problem first asks for the theoretical yield of PbI_2 when two quantities of reactants are mixed. Before we can calculate the amount of product we need to know which reactant is limiting. Use the molar mass for each product and the mole ratio for the balanced equation to calculate the moles of PbI_2 that could be produced.

$$125.5 \text{ g NaI} \times \frac{1 \text{ mol NaI}}{149.89 \text{ g NaI}} \times \frac{1 \text{ mol } PbI_2}{2 \text{ mol NaI}} = 0.4186 \text{ mol } PbI_2$$

$$205.6 \text{ g } Pb(NO_3)_2 \times \frac{1 \text{ mol } Pb(NO_3)_2}{331.22 \text{ g } Pb(NO_3)_2} \times \frac{1 \text{ mol } PbI_2}{1 \text{ mol } Pb(NO_3)_2} = 0.6207 \text{ mol } PbI_2$$

The limiting reactant is NaI. Now we can answer the question about theoretical yield. Theoretical is the amount of product we calculate can be produced, that is, 0.4186 mol PbI_2. In real life, the actual yield might be less than the calculated yield. The theoretical yield of PbI_2 can be calculated from the number of moles of PbI_2, if we know the molar mass.

$$0.4186 \text{ mol } PbI_2 \times \frac{461.00 \text{ g } PbI_2}{1 \text{ mol } PbI_2} = 193.0 \text{ g } PbI_2$$

b. In part a we calculated the theoretical yield of lead(II) iodide, which is 193.0 g. We are told that the actual yield from this reaction was found to be 164.5 g. The percent yield is equal to the actual yield divided by the theoretical yield, multiplied by 100 percent. So the percent yield of lead(II) iodide is

$$\frac{164.5 \text{ g } PbI_2}{193.0 \text{ g } PbI_2} \times 100\% = 85.23\%$$