QUANTITIES IN CHEMICAL REACTIONS

Chapter Eight

Conservation

- Recall that, in any chemical reaction, that matter is always conserved.
- Consider the reaction of zinc with sulfur:

$$Zn_{(s)} + S_{(s)} \rightarrow ZnS_{(s)}$$

- From this reaction we can imply that, assuming the reaction works perfectly, every one atom of zinc reacts with one atom of sulfur to produce one unit of zinc sulfide.
- More conveniently, we can say that one mole of zinc atoms (1 mol Zn) reacts with one mole of sulfur atoms (1 mol S), producing one mole of zinc sulfide (1 mol ZnS)

Common Misconceptions

- Is it true that if I start a reaction with 5.0 g of zinc and 5.0 g of sulfur that I will end the reaction with 10.0 g of *matter*?
- Is it true that if I start a reaction with 5.0 g of zinc and 5.0 g of sulfur that I will end the reaction with 10.0 g of *zinc sulfide*?
- In predicting the quantities of products created during a reaction, it is ultimately the number of atoms/molecules reacting that we must consider, not the mass.
- The most convenient way to do this is to consider the moles of reactants participating in a reaction.

Coefficients of Reactions

- The coefficients in chemical equations are useful in this context because they tell us how many moles of reactants are required to produce a certain number of moles of products.
- We can say that one mole of zinc reacts with one mole of sulfur, producing 1 mole of zinc sulfide. The three are equivalent for the purpose of this reaction.

$$1 \text{ mol Zn} \leftrightarrow 1 \text{mol S} \leftrightarrow 1 \text{ mol ZnS}$$

Coefficients of Reactions

- We can convert between moles of reactants and products using dimensional analysis, the same technique we used for unit conversions.
- To convert between moles of zinc and moles of zinc sulfide, we use the following ratio:

$$\frac{1\,\text{mol ZnS}}{1\,\text{mol Zn}}$$

- Other similar ratios between zinc and sulfur, and between sulfur and zinc sulfide can be created.

Basic Problems

Let's consider the following reaction:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$$

To produce 2 mol of water we need to combine_____mol H_2 with_____mol O_2.

How many moles of each reactant are required to produce 10. moles of water?

How many moles of water can be produced at most from 6.0 moles of hydrogen?

How many moles of oxygen would I need to accomplish this?

Basic Problems

In the reaction of aqueous solutions of hydrochloric acid and barium hydroxide:

a. How many moles of each reactant do you need to produce 5.75 mol water?
b. What is the maximum number of moles of water you can produce with 18.0 mol HCl? How many moles of barium hydroxide would you need to accomplish this?

Using the Mass-Mole Relationship

- Recall that the molar mass relates the mass of a compound (or element) to the number of moles.
- Therefore, if we know the mass of a given reactant, we can easily determine the number of moles of the reactant, and from there the number of moles of product produced.
- Remember, you must use the mole-mole relationship to carry out a conversion; the mass alone is not sufficient!

Mass-Mole Problems

Consider the reaction of sodium metal with chlorine to make sodium chloride.

a. How many moles of sodium are contained in 15.50 g of sodium?

b. Assuming we have excess (that is, more than enough) chlorine, how many moles of sodium chloride could be produced?

c. How many grams of sodium chloride is this?

Mass-Mole Problems

When 39.75 g of sodium is reacted with excess chlorine, how many grams of sodium chloride can be produced?

When 146.7 g of chlorine is reacted with excess sodium, how many grams of sodium chloride can be produced?

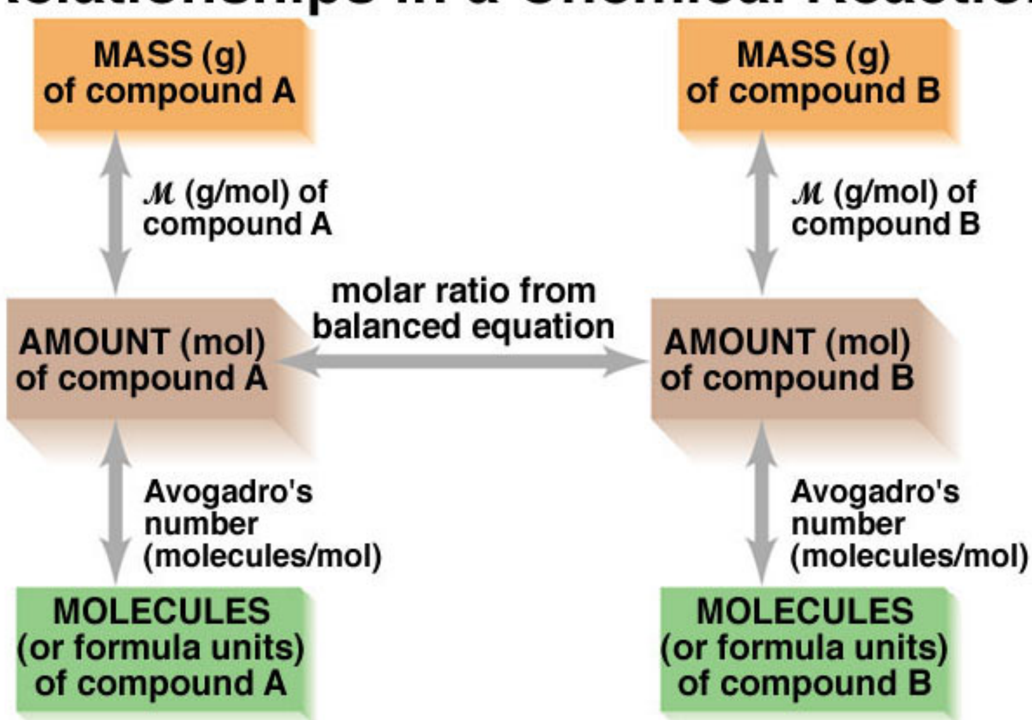
Mass-Mole Problems

Ethanol has molecular formula CH_3CH_2OH.

a. How many grams of carbon dioxide and water can be produced if 55.0 mL of ethanol is combusted?
Note: The density of ethanol is 0.789 g/mL.

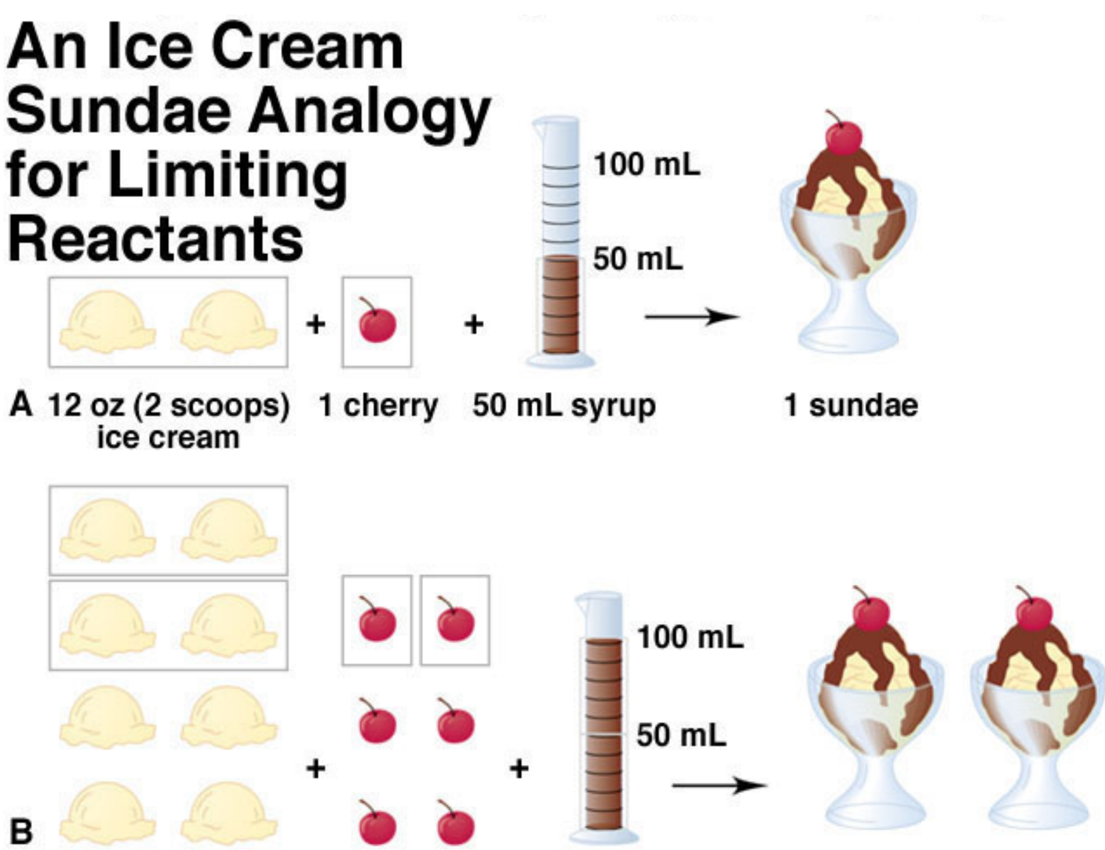
b. How many oxygen molecules are required to react with this amount of ethanol?

Summary of the Mass-Mole-Number Relationships in a Chemical Reaction



The Ice Cream Sundae Problem

- Suppose we are going to make ice cream sundaes.
- We decide that each of our sundaes must contain
  - 2 scoops of ice cream
  - 50 mL of chocolate syrup
  - 1 cherry
- How many sundaes can I make if I have 8 scoops of ice cream, 13 cherries, and 300 mL of chocolate syrup available?



Limiting Reactants

- The problem we had with our ice cream sundaes is the same type of problem we face in chemical reactions.
- In virtually all reactions, we will have a disproportionate amount of reactants.
- The reactant which will produce *less* product is called the limiting reactant; in all cases, it predicts the amount of product formed.
- The other reactants are said to be in excess, meaning there is more than enough of it.
- The key to identifying a limiting reagent problem is that it will give you the amounts being reacted of more than one reactant.

Limiting Reactants

- Suppose you with to synthesize zinc sulfide.
- You combine 50. grams of zinc with 50. grams of sulfur.
- What mass of zinc sulfide is predicted to be produced from the zinc?
- What mass of zinc sulfide is predicted to be produced from the sulfur?
- Which reactant is the limiting reactant?

Examples

What mass of water is produced by the reaction of 75.8 g of oxygen with 10.5 g of hydrogen?

Examples

Consider the bromination of propyne(C_3H_4):

$$C_3H_4 + 2\ Br_2 \rightarrow C_3H_4Br_4$$

78.24 g of propyne is treated with 125.0 g of bromine.

a. What mass of product ($C_3H_4Br_4$) is produced?

b. Which reactant was in excess, and what mass of it remains after the reaction is completed?

Percent Yield

- So far, we have assumed that all reactions will produce exactly the amount of products that our calculations will predict.
- The *theoretical yield* is the amount of a product you expect to get at the end of a reaction.
- In reality, we never actually get 100% of this amount.
  - Why might this be so?
- The amount of a product that is *really* produced at the end of a reaction is called the *actual yield*.

Percent Yield

- The *percent yield* gives us an idea of how much of a product we actually produced compared to the amount that should have been produced.

$$\%\,\text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

- In general, the closer the percent yield is to 100% percent, the better the reaction

Example

- 8.0 g of hydrogen is combined with excess oxygen, producing 58.5 g of water. What is the percent yield of this reaction?

Example

Suppose that you wish to produce 15.00 grams of barium sulfate by reacting excess sulfuric acid with barium hydroxide. If this reaction is known to usually give an 87.5% yield, how many grams of barium hydroxide should you use?