CHAPTER 9: STOICHIOMETRY USING CHEMICAL EQUATIONS

Enhanced Introductory College Chemistry

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Chapter Contents

Please visit the web version of Enhanced Introductory College Chemistry (https://ecampusontario.pressbooks.pub/enhancedchemistry/) to access the complete book, interactive activities and ancillary resources.

In this chapter, you will learn about

- Quantitative components within a chemical reaction
- Mole ratios within a chemical reaction

- Stoichiometry analysis of a chemical reaction involving moles and mass
- Determine the limiting reactant and the excess reactant in a chemical reaction
- Determine the percent yield within a chemical reaction

To better support your learning, you should be familiar with the following concepts before starting this chapter:

- Nomenclature
- Compounds
- Mole quantities and calculations
- Molar mass

Figure 9a Example of a chemistry building that required stoichiometric quantifications to be constructed. (credit: Photo by Bree Evans, Unsplash license)

At Contrived State University in Anytown, Ohio, a new building was dedicated in March 2010 to house the College of Education. The 100,000-square-foot building has enough office space to accommodate 86 full-time faculty members and 167 full-time staff.

In a fit of monetary excess, the university administration offered to buy new furniture (desks and chairs) and computer workstations for all faculty and staff members moving into the new building. However, to save on long-term energy and materials costs, the university offered to buy only 1 laser printer per 10 employees, with the plan to network the printers together.

How many laser printers did the administration have to buy? It is rather simple to show that 26 laser printers are needed for all the employees. However, what if a chemist was calculating quantities for a chemical reaction? Interestingly enough, similar calculations can be performed for chemicals as well as laser printers.

In filling a new office building with furniture and equipment, managers do calculations similar to those performed by scientists doing chemical reactions.

We have already established that quantities are important in science, especially in chemistry. It is important to make accurate measurements of a variety of quantities when performing experiments. However, it is also important to be able to relate one measured quantity to another, unmeasured quantity. In this chapter, we will consider how we manipulate quantities to relate them to each other.

Attribution & References

9.1 STOICHIOMETRY BASICS

Learning Objectives

By the end of this section, you will be able to:

- Explain the concept of stoichiometry as it pertains to chemical reactions
- Use balanced chemical equations to derive stoichiometric factors relating amounts of reactants and products

A balanced chemical equation provides a great deal of information in a very succinct format. Chemical formulas provide the identities of the reactants and products involved in the chemical change, allowing classification of the reaction. Coefficients provide the relative numbers of these chemical species, allowing a quantitative assessment of the relationships between the amounts of substances consumed and produced by the reaction. These quantitative relationships are known as the reaction's **stoichiometry**, a term derived from the Greek words *stoicheion* (meaning "element") and *metron* (meaning "measure"). In this module, the use of balanced chemical equations for various stoichiometric applications is explored.

The general approach to using stoichiometric relationships is similar in concept to the way people go about many common activities. Food preparation, for example, offers an appropriate comparison. A recipe for making eight pancakes calls for 1 cup pancake mix, $\frac{3}{4}$ cup milk, and one egg. The "equation" representing the preparation of pancakes per this recipe is

$$1 \text{ cup mix} + \frac{3}{4} \text{ cup milk} + 1 \text{ egg} \longrightarrow 8 \text{ pancakes}$$

If two dozen pancakes are needed for a big family breakfast, the ingredient amounts must be increased proportionally according to the amounts given in the recipe. For example, the number of eggs required to make 24 pancakes is

$$24 \cancel{\text{pancakes}} \times \frac{1 \text{ egg}}{8 \cancel{\text{pancakes}}} = 3 \text{ eggs}$$

Balanced chemical equations are used in much the same fashion to determine the amount of one reactant required to react with a given amount of another reactant, or to yield a given amount of product, and so forth. The coefficients in the balanced equation are used to derive **stoichiometric factors** that permit computation of the desired quantity. To illustrate this idea, consider the production of ammonia by reaction of hydrogen and nitrogen:

$$\mathrm{N_2}(g) + \mathrm{3H_2}(g) \longrightarrow \mathrm{2NH_3}(g)$$

This equation shows ammonia molecules are produced from hydrogen molecules in a 2:3 ratio, and stoichiometric factors may be derived using any amount (number) unit:

$$\frac{2\ \mathrm{NH_3\ molecules}}{3\ \mathrm{H_2\ molecules}}\ \text{or}\ \frac{2\ \mathrm{doz\ NH_3\ molecules}}{3\ \mathrm{doz\ H_2\ molecules}}\ \text{or}\ \frac{2\ \mathrm{mol\ NH_3\ molecules}}{3\ \mathrm{mol\ H_2\ molecules}}$$

These stoichiometric factors can be used to compute the number of ammonia molecules produced from a given number of hydrogen molecules, or the number of hydrogen molecules required to produce a given number of ammonia molecules. Similar factors may be derived for any pair of substances in any chemical equation.

Watch Determining the Mole Ratio (6 mins) (https://www.youtube.com/watch?v=42Mk1B2u0hk)

Example 9.1a

Moles of Reactant Required in a Reaction

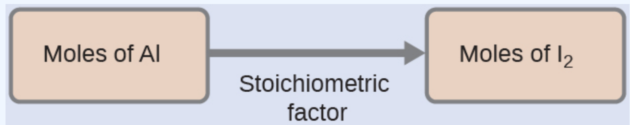
How many moles of I_2 are required to react with 0.429 mol of Al according to the following equation (see Figure 9.1a)?

$$\mathrm{2Al} + \mathrm{3I_2} \longrightarrow \mathrm{2AlI_3}$$

Figure 9.1a Aluminum and iodine react to produce aluminum iodide. The heat of the reaction vaporizes some of the solid iodine as a purple vapour. (credit: modification of work by Mark Ott in *Chemistry (OpenStax)*, CC BY 4.0).

Solution

Referring to the balanced chemical equation, the stoichiometric factor relating the two substances of interest is $\frac{3\ \mathrm{mol\ I_2}}{2\ \mathrm{mol\ Al}}$. The molar amount of iodine is derived by multiplying the provided molar amount of aluminum by this factor:



$$0.429\ \cancel{\text{mol Al}} \times \frac{3\ \text{mol I}_2}{2\ \cancel{\text{mol Al}}}$$

$$= 0.644\ \text{mol I}_2$$

Exercise 9.1a

How many moles of $Ca(OH)_2$ are required to react with 1.36 mol of H_3PO_4 to produce $Ca_3(PO_4)_2$ according to the equation:

$$3\text{Ca(OH)}_2 + 2\text{H}_3\text{PO}_4 \longrightarrow \text{Ca}_3(\text{PO}_4)_2 + 6\text{H}_2\text{O}?$$

Check Your Answer[1]

These examples illustrate the ease with which the amounts of substances involved in a chemical reaction of known stoichiometry may be related.

Exercise 9.1b

Practice using the following PhET simulation: Reactants, Products, and Leftovers (https://phet.colorado.edu/sims/html/reactants-products-and-leftovers/latest/reactants-products-and-leftovers_en.html)

Link to Interactive Learning Tools

Practice Mole-to-mole ratios (https://h5pstudio.ecampusontario.ca/content/21799) from eCampusOntario H5P Studio (https://h5pstudio.ecampusontario.ca/).

Practice Stoichiometry – Relationships (https://www.physicsclassroom.com/Concept-Builders/Chemistry/Stoichiometry-Table) (Apprentice Difficulty Level ONLY) from The Physics Classroom (https://www.physicsclassroom.com/).

Attribution & References

Notes

1. 2.04 mol $Ca(OH)_2$

9.2 MOLE-MASS AND MASS-MASS CALCULATIONS

Learning Objectives

By the end of this section, you will be able to:

- From a given number of moles of a substance, calculate the mass of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the moles of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the mass of another substance involved using the balanced chemical equation.

Mole-mole calculations are not the only type of calculations that can be performed using balanced chemical equations. Recall that the molar mass can be determined from a chemical formula and used as a conversion factor. We can add that conversion factor as another step in a calculation to make a **mole-mass calculation**, where we start with a given number of moles of a substance and calculate the mass of another substance involved in the chemical equation, or vice versa.

For example, suppose we have the balanced chemical equation:

$$2Al + 3Cl_2 \rightarrow 2AlCl_3$$

Suppose we know we have 123.2 g of Cl_2. How can we determine how many moles of $AlCl_3$ we will get when the reaction is complete? First and foremost, *chemical equations are not balanced in terms of grams; they are balanced in terms of moles*. So to use the balanced chemical equation to relate an amount of Cl_2 to an amount of $AlCl_3$, we need to convert the given amount of Cl_2 into moles. We know how to do this by simply using the molar mass of Cl_2 as a conversion factor. The molar mass of Cl_2 (which we get from the atomic mass of Cl from the periodic table) is 70.90 g/mol. We must invert this fraction so that the units cancel properly:

$$123.2\ \cancel{g\ Cl_2} \times \frac{1\ mol\ Cl_2}{70.90\ \cancel{g\ Cl_2}} = 1.738\ mol\ Cl_2$$

Now that we have the quantity in moles, we can use the balanced chemical equation to construct a conversion factor that relates the number of moles of Cl_2 to the number of moles of $AlCl_3$. The numbers in the conversion factor come from the coefficients in the balanced chemical equation:

$$\frac{2\ mol\ AlCl_3}{3\ mol\ Cl_2}$$

Using this conversion factor with the molar quantity we calculated above, we get:

$$1.738\ \cancel{mol\ Cl_2} \times \frac{2\ mol\ AlCl_3}{3\ \cancel{mol\ Cl_2}} = 1.159\ mol\ AlCl_3$$

So, we will get 1.159 mol of $AlCl_3$ if we react 123.2 g of Cl_2.

In this last example, we did the calculation in two steps. However, it is mathematically equivalent to perform the two calculations sequentially on one line:

$$123.2\ \cancel{g\ Cl_2} \times \frac{1\ \cancel{mol\ Cl_2}}{70.90\ \cancel{g\ Cl_2}} \times \frac{2\ mol\ AlCl_3}{3\ \cancel{mol\ Cl_2}} = 1.159\ mol\ AlCl_3$$

The units still cancel appropriately, and we get the same numerical answer in the end. Sometimes the answer may be slightly different from doing it one step at a time because of rounding of the intermediate answers, but the final answers should be effectively the same.

Example 9.2a

Problem

How many moles of HCl will be produced when 249 g of $AlCl_3$ are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Solution

We will do this in two steps: convert the mass of $AlCl_3$ to moles and then use the balanced chemical equation to find the number of moles of HCl formed. The molar mass of $AlCl_3$ is 133.33 g/mol, which we have to invert to get the appropriate conversion factor:

$$1.87\ \cancel{mol\ AlCl_3} \times \frac{6\ mol\ HCl}{2\ \cancel{mol\ AlCl_3}} = 5.61\ mol\ HCl$$

Now we can use this quantity to determine the number of moles of HCl that will form. From the balanced chemical equation, we construct a conversion factor between the number of moles of $AlCl_3$ and the number of moles of HCl:

$$\frac{6\ \text{mol HCl}}{2\ \text{mol AlCl}_3}$$

Applying this conversion factor to the quantity of $AlCl_3$, we get:

$$1.87\ \cancel{\text{mol AlCl}_3} \times \frac{6\ \text{mol HCl}}{2\ \cancel{\text{mol AlCl}_3}} = 5.61\ \text{mol HCl}$$

Alternatively, we could have done this in one line:

$$249\ \cancel{\text{g AlCl}_3} \times \frac{1\ \cancel{\text{mol AlCl}_3}}{133.33\ \cancel{\text{g AlCl}_3}} \times \frac{6\ \text{mol HCl}}{2\ \cancel{\text{mol AlCl}_3}} = 5.60\ \text{mol HCl}$$

The last digit in our final answer is slightly different because of rounding differences, but the answer is essentially the same.

Exercise 9.2a

How many moles of Al_2O_3 will be produced when 23.9 g of H_2O are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Check Your Answer[1]

A variation of the mole-mass calculation is to start with an amount in moles and then determine an amount of another substance in grams. The steps are the same but are performed in reverse order.

Example 9.2b

Problem

How many grams of NH_3 will be produced when 33.9 mol of H_2 are reacted according to this chemical equation?

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Solution

The conversions are the same, but they are applied in a different order. Start by using the balanced chemical equation to convert to moles of another substance and then use its molar mass to determine the mass of the final substance. In two steps, we have:

$$33.9\ \cancel{\text{mol } H_2} \times \frac{2\text{ mol } NH_3}{3\ \cancel{\text{mol } H_2}} = 22.6\text{ mol } NH_3$$

Now, using the molar mass of NH_3, which is 17.03 g/mol, we get:

$$22.6\ \cancel{\text{mol } NH_3} \times \frac{17.03\text{ g } NH_3}{1\ \cancel{\text{mol } NH_3}} = 385\text{ g } NH_3$$

Exercise 9.2b

How many grams of N_2 are needed to produce 2.17 mol of NH_3 when reacted according to this chemical equation?

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Check Your Answer[2]

It should be a trivial task now to extend the calculations to **mass-mass calculations**, in which we start with a mass of some substance and end with the mass of another substance in the chemical reaction. For this type of calculation, the molar masses of two different substances must be used—be sure to keep track of which is which. Again, however, it is important to emphasize that before the balanced chemical reaction is used, the mass quantity must first be converted to moles. Then the coefficients of the balanced chemical reaction can be used to convert to moles of another substance, which can then be converted to a mass.

For example, let us determine the number of grams of SO_3 that can be produced by the reaction of 45.3 g of SO_2 and O_2:

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

First, we convert the given amount, 45.3 g of SO_2, to moles of SO_2 using its molar mass (64.06 g/mol):

$$45.3\ \cancel{\text{g SO}_2} \times \frac{1\ \text{mol SO}_2}{64.06\ \cancel{\text{g SO}_2}} = 0.707\ \text{mol SO}_2$$

Second, we use the balanced chemical reaction to convert from moles of SO_2 to moles of SO_3:

$$0.707\ \cancel{\text{mol SO}_2} \times \frac{2\ \text{mol SO}_3}{2\ \cancel{\text{mol SO}_2}} = 0.707\ \text{mol SO}_3$$

Finally, we use the molar mass of SO_3 (80.06 g/mol) to convert to the mass of SO_3:

$$0.707\ \cancel{\text{mol SO}_3} \times \frac{80.06\ \text{g SO}_3}{1\ \cancel{\text{mol SO}_3}} = 56.6\ \text{g SO}_3$$

We can also perform all three steps sequentially, writing them on one line as:

$$45.3\ \cancel{\text{g SO}_2} \times \frac{1\ \cancel{\text{mol SO}_2}}{64.06\ \cancel{\text{g SO}_2}} \times \frac{2\ \cancel{\text{mol SO}_3}}{2\ \cancel{\text{mol SO}_2}} \times \frac{80.06\ \text{g SO}_3}{1\ \cancel{\text{mol SO}_3}} = 56.6\ \text{g SO}_3$$

We get the same answer. Note how the initial and all the intermediate units cancel, leaving grams of SO_3, which is what we are looking for, as our final answer.

Example 9.2c

Problem

What mass of Mg will be produced when 86.4 g of K are reacted?

$$MgCl_2(s) + 2K(s) \rightarrow Mg(s) + 2KCl(s)$$

Solution

We will simply follow the steps:

mass K → mol K → mol Mg → mass Mg

In addition to the balanced chemical equation, we need the molar masses of K (39.09 g/mol) and Mg (24.31 g/mol). In one line,

$$86.4\ \cancel{\text{g K}} \times \frac{1\ \cancel{\text{mol K}}}{39.09\ \cancel{\text{g K}}} \times \frac{1\ \cancel{\text{mol Mg}}}{2\ \cancel{\text{mol K}}} \times \frac{24.31\ \text{g Mg}}{1\ \cancel{\text{mol Mg}}} = 26.87\ \text{Mg}$$

Exercise 9.2c

What mass of H_2 will be produced when 122 g of Zn are reacted?

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

Check Your Answer[3]

Example 9.2d

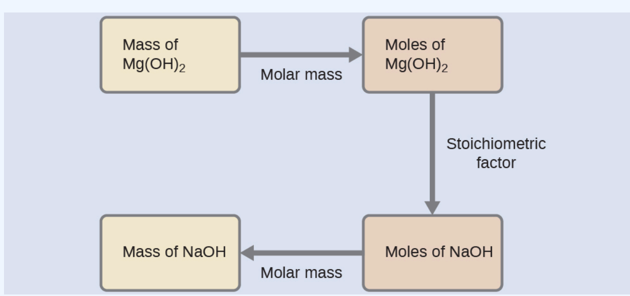
Relating Masses of Reactants and Products

What mass of sodium hydroxide, NaOH, would be required to produce 16 g of the antacid milk of magnesia [magnesium hydroxide, $Mg(OH)_2$] by the following reaction?

$$\mathrm{MgCl_2}(aq) + 2\mathrm{NaOH}(aq) \longrightarrow \mathrm{Mg(OH)_2}(s) + \mathrm{NaCl}(aq)$$

Solution

The approach used previously in Example 1 and Example 2 is likewise used here; that is, we must derive an appropriate stoichiometric factor from the balanced chemical equation and use it to relate the amounts of the two substances of interest. In this case, however, masses (not molar amounts) are provided and requested, so additional steps of the sort learned in the previous chapter are required. The calculations required are outlined in this flowchart:



$$16\ \cancel{\mathrm{g\ Mg(OH)_2}} \times \frac{1\ \cancel{\mathrm{mol\ Mg(OH)_2}}}{58.3\ \cancel{\mathrm{g\ Mg(OH)_2}}} \times \frac{2\ \cancel{\mathrm{mol\ NaOH}}}{1\ \cancel{\mathrm{mol\ Mg(OH)_2}}} \times \frac{40.0\ \mathrm{g\ NaOH}}{1\ \cancel{\mathrm{mol\ NaOH}}} = 22\ \mathrm{g\ NaOH}$$

Exercise 9.2d

What mass of gallium oxide, Ga_2O_3, can be prepared from 29.0 g of gallium metal? The equation for the reaction is:

$$4Ga + 3O_2 \longrightarrow 2Ga_2O_3$$

Check Your Answer[4]

Example 9.2e

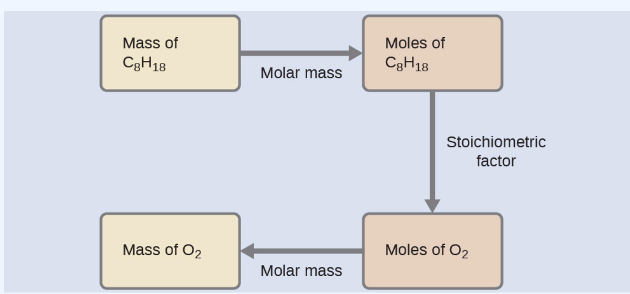
Relating Masses of Reactants

What mass of oxygen gas, O_2, from the air is consumed in the combustion of 702 g of octane, C_8H_{18}, one of the principal components of gasoline?

$$2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$$

Solution

The approach required here is the same as for the Example 3, differing only in that the provided and requested masses are both for reactant species.



$$702\ \cancel{\text{g C}_8\text{H}_{18}} \times \frac{1\ \cancel{\text{mol C}_8\text{H}_{18}}}{114.23\ \cancel{\text{g C}_8\text{H}_{18}}} \times \frac{25\ \cancel{\text{mol O}_2}}{2\ \cancel{\text{mol C}_8\text{H}_{18}}} \times \frac{32.00\ \text{g O}_2}{\cancel{\text{mol O}_2}} = 2.46 \times 10^3\ \text{g O}_2$$

Exercise 9.2e

What mass of CO is required to react with 25.13 g of Fe_2O_3 according to the equation:

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

Check Your Answer[5]

These examples illustrate just a few instances of reaction stoichiometry calculations. Numerous variations on the beginning and ending computational steps are possible depending upon what particular quantities are provided and sought (volumes, solution concentrations, and so forth). Regardless of the details, all these calculations share a common essential component: the use of stoichiometric factors derived from balanced chemical equations. Figure 9.2a provides a general outline of the various computational steps associated with many reaction stoichiometry calculations.

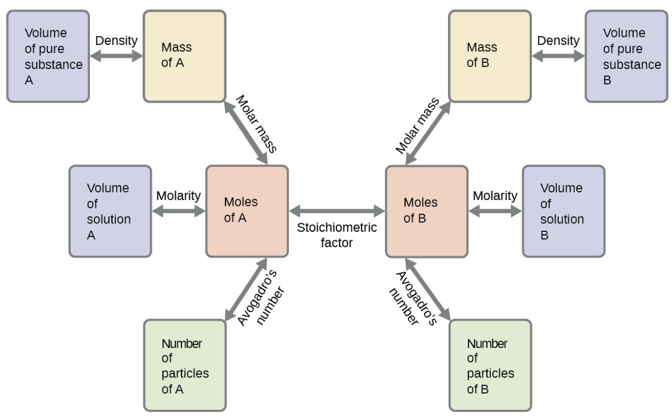


Figure 9.2a The flowchart depicts the various computational steps involved in most reaction stoichiometry calculations (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Airbags

Airbags (Figure 9.2b) are a safety feature provided in most automobiles since the 1990s. The effective operation of an airbag requires that it be rapidly inflated with an appropriate amount (volume) of gas when the vehicle is involved in a collision. This requirement is satisfied in many automotive airbag systems through use of explosive chemical reactions, one common choice being the decomposition of sodium azide, NaN_3. When sensors in the vehicle detect a collision, an electrical current is passed through a carefully measured amount of NaN_3 to initiate its decomposition:

$$2NaN_3(s) \longrightarrow 3N_2(g) + 2Na(s)$$

This reaction is very rapid, generating gaseous nitrogen that can deploy and fully inflate a typical airbag in a fraction of a second (~0.03–0.1 s). Among many engineering considerations, the amount of sodium azide used must be appropriate for generating enough nitrogen gas to fully inflate the air bag and ensure its proper function. For example, a small mass (~100 g) of NaN_3 will generate approximately 50 L of N_2.

Figure 9.2b Airbags deploy upon impact to minimize serious injuries to passengers. (credit: work by Jon Seidman, CC BY 2.0)

Links to Interactive Learning Tools

Practice Stoichiometry – Relationships (https://www.physicsclassroom.com/Concept-Builders/Chemistry/Stoichiometry-Table) (all levels) from The Physics Classroom (https://www.physicsclassroom.com/).

Attribution & References

Except where otherwise noted, this page is adapted by Adrienne Richards from:

Notes

1. 0.442 mol
2. 30.4 g (Note: here we go from a product to a reactant, showing that mole-mass problems can begin and end with any substance in the chemical equation.)
3. 3.77 g
4. 39.0 g Ga_2O_3
5. 13

9.3 LIMITING REACTANTS

Learning Objectives

By the end of this section, you will be able to:

- Explain the concepts of limiting reactants/reagents.

The relative amounts of reactants and products represented in a balanced chemical equation are often referred to as *stoichiometric amounts*. All the exercises of the preceding module involved stoichiometric amounts of reactants. For example, when calculating the amount of product generated from a given amount of reactant, it was assumed that any other reactants required were available in stoichiometric amounts (or greater). In this module, more realistic situations are considered, in which reactants are not present in stoichiometric amounts.

Limiting Reactant

Consider another food analogy, making grilled cheese sandwiches (Figure 9.2a):

$$\text{1 slice of cheese} + \text{2 slices of bread} \longrightarrow \text{1 sandwich}$$

Stoichiometric amounts of sandwich ingredients for this recipe are bread and cheese slices in a 2:1 ratio. Provided with 28 slices of bread and 11 slices of cheese, one may prepare 11 sandwiches per the provided recipe, using all the provided cheese and having six slices of bread left over. In this scenario, the number of sandwiches prepared has been *limited* by the number of cheese slices, and the bread slices have been provided in *excess*.

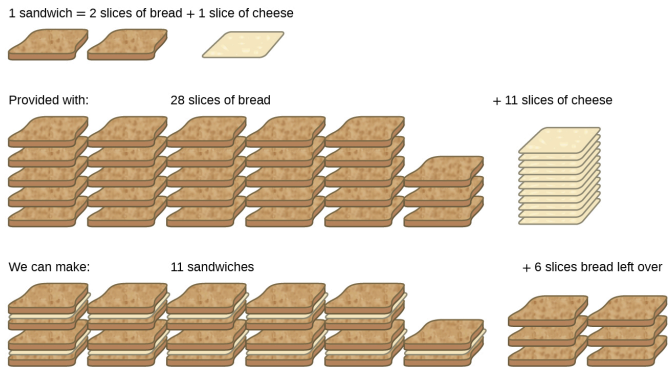


Figure 9.3a Sandwich making can illustrate the concepts of limiting and excess reactants (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Consider this concept now with regard to a chemical process, the reaction of hydrogen with chlorine to yield hydrogen chloride:

$$H_2(s) + Cl_2(g) \longrightarrow 2HCl(g)$$

The balanced equation shows the hydrogen and chlorine react in a 1:1 stoichiometric ratio. If these reactants are provided in any other amounts, one of the reactants will nearly always be entirely consumed, thus limiting the amount of product that may be generated. This substance is the **limiting reactant**, and the other substance is the **excess reactant**. Identifying the limiting and excess reactants for a given situation requires computing the molar amounts of each reactant provided and comparing them to the stoichiometric amounts represented in the balanced chemical equation. For example, imagine combining 3 moles of H_2 and 2 moles of Cl_2. This represents a 3:2 (or 1.5:1) ratio of hydrogen to chlorine present for reaction, which is greater than the stoichiometric ratio of 1:1. Hydrogen, therefore, is present in excess, and chlorine is the limiting reactant. Reaction of all the provided chlorine (2 mol) will consume 2 mol of the 3 mol of hydrogen provided, leaving 1 mol of hydrogen unreacted.

An alternative approach to identifying the limiting reactant involves comparing the amount of product expected for the complete reaction of each reactant. Each reactant amount is used to separately calculate the amount of product that would be formed per the reaction's stoichiometry. The reactant yielding the lesser amount of product is the limiting reactant. For the example in the previous paragraph, complete reaction of the hydrogen would yield

$$\text{mol HCl produced} = 3 \text{ mol } H_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } H_2} = 6 \text{ mol HCl}$$

Complete reaction of the provided chlorine would produce

$$\text{mol HCl produced} = 2 \text{ mol } Cl_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } Cl_2} = 4 \text{ mol HCl}$$

The chlorine will be completely consumed once 4 moles of HCl have been produced. Since enough hydrogen was provided to yield 6 moles of HCl, there will be unreacted hydrogen remaining once this reaction is complete. Chlorine, therefore, is the limiting reactant and hydrogen is the excess reactant (Figure 9.2b).

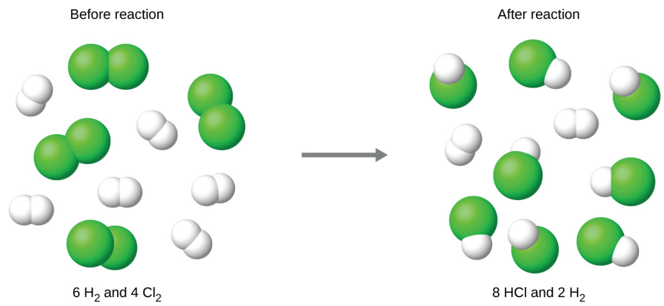


Figure 9.3b When H_2 and Cl_2 are combined in nonstoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant (credit: *Chemistry (OpenStax),* CC BY 4.0).

Exercise 9.3a

Practice using the following PhET simulation: Reactants, Products, and Leftovers (https://phet.colorado.edu/sims/html/reactants-products-and-leftovers/latest/reactants-products-and-leftovers_en.html)

Example 9.3a

Identifying the Limiting Reactant and Calculate Max Amount of Product Produced

Silicon nitride is a very hard, high-temperature-resistant ceramic used as a component of turbine blades in jet engines. It is prepared according to the following equation:

$$3\text{Si}(s) + 2\text{N}_2(g) \longrightarrow \text{Si}_3\text{N}_4(s)$$

A) Which is the limiting reactant when 2.00 g of Si and 1.50 g of N_2 react?

B) How many grams of silicon nitride are produced?

Solution

A) Compute the provided molar amounts of reactants, and then compare these amounts to the balanced equation to identify the limiting reactant.

$$\text{mol Si} = 2.00\ \cancel{\text{g Si}} \times \frac{1\ \text{mol Si}}{28.09\ \cancel{\text{g Si}}} = 0.0712\ \text{mol Si}$$

$$\text{mol N}_2 = 1.50\ \cancel{\text{g N}_2} \times \frac{1\ \text{mol N}_2}{28.09\ \cancel{\text{g N}_2}} = 0.0535\ \text{mol N}_2$$

The provided Si:N_2 molar ratio is:

$$\frac{0.0712\ \text{mol Si}}{0.0535\ \text{mol N}_2} = \frac{1.33\ \text{mol Si}}{1\text{mol N}_2}$$

The stoichiometric Si:N_2 ratio is:

$$\frac{3\ \text{mol Si}}{2\ \text{mol N}_2} = \frac{1.5\ \text{mol Si}}{1\ \text{mol N}_2}$$

Comparing these ratios shows that Si is provided in a less-than-stoichiometric amount, and so is the limiting reactant.

Alternatively, compute the amount of product expected for complete reaction of each of the provided reactants. The 0.0712 moles of silicon would yield

$$\text{mol Si}_3\text{N}_4\ \text{produced} = 0.0712\ \text{mol Si} \times \frac{1\ \text{mol Si}_3\text{N}_4}{3\ \text{mol Si}} = 0.0237\ \text{mol Si}_3\text{N}_4$$

while the 0.0535 moles of nitrogen would produce

$$\text{mol Si}_3\text{N}_4\ \text{produced} = 0.0535\ \text{mol N}_2 \times \frac{1\ \text{mol Si}_3\text{N}_4}{2\ \text{mol N}_2} = 0.0268\ \text{mol Si}_3\text{N}_4$$

Since silicon yields the lesser amount of product, it is the limiting reactant.

B) In order to determine the maximum amount of product produced, first the limiting reactant has to be determined. In part A, the limiting reactant was determined to be silicon. Therefore, the moles of silicon nitride produced by silicon will be used to calculate how many grams of silicon nitride is produced.

$$0.0237\ \text{mol Si}_3\text{N}_4 \times \frac{140.28\ \text{g Si}_3\text{N}_4}{1\ \text{mol Si}_3\text{N}_4} = 3.32\ \text{grams Si}_3\text{N}_4$$

Exercise 9.3b

Which is the limiting reactant when 5.00 g of H_2 and 10.0 g of O_2 react and form water? How many grams of water is produced?

Check Your Answer[1]

Indigenous Perspective: Science of Bannock

Bannock is a staple food in many Indigenous communities throughout Canada. Here is an Anishinaabeg example of the science of Bannock.

Watch The Science of Bannock (23 mins) (https://www.youtube.com/watch?v=m02hd32QJ8I)

Attribution & References

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Notes

1. O_2 is limiting reactant, 22.5 g H_2O produced

9.4 REACTION YIELDS

Learning Objectives

By the end of this section, you will be able to:

- Explain the concepts of theoretical yield.
- Derive the theoretical yield for a reaction under specified conditions.
- Calculate the percent yield for a reaction.

Percent Yield

The amount of product that *may be* produced by a reaction under specified conditions, as calculated per the stoichiometry of an appropriate balanced chemical equation, is called the **theoretical yield** of the reaction. In practice, the amount of product obtained is called the **actual yield**, and it is often less than the theoretical yield for a number of reasons. Some reactions are inherently inefficient, being accompanied by *side reactions* that generate other products. Others are, by nature, incomplete (consider the partial reactions of weak acids and bases discussed earlier in this chapter). Some products are difficult to collect without some loss, and so less than perfect recovery will reduce the actual yield. The extent to which a reaction's theoretical yield is achieved is commonly expressed as its **percent yield**:

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

Actual and theoretical yields may be expressed as masses or molar amounts (or any other appropriate property; e.g., volume, if the product is a gas). As long as both yields are expressed using the same units, these units will cancel when percent yield is calculated.

Example 9.4a

Calculation of Percent Yield

Upon reaction of 1.274 g of copper sulfate with excess zinc metal, 0.392 g copper metal was obtained according to the equation:

$$\mathrm{CuSO_4}(aq) + \mathrm{Zn}(s) \longrightarrow \mathrm{Cu}(s) + \mathrm{ZnSO_4}(aq)$$

What is the percent yield?

Solution

The provided information identifies copper sulfate as the limiting reactant, and so the theoretical yield is found by the approach illustrated in the previous module, as shown here:

$$1.274\ \cancel{\mathrm{g\ CuSO_4}} \times \frac{1\ \cancel{\mathrm{mol\ CuSO_4}}}{159.62\ \cancel{\mathrm{g\ CuSO_4}}} \times \frac{1\ \cancel{\mathrm{mol\ Cu}}}{1\cancel{\mathrm{mol\ CuSO_4}}} \times \frac{63.55\ \mathrm{g\ Cu}}{1\ \cancel{\mathrm{mol\ Cu}}} = 0.5072\ \mathrm{g\ Cu}$$

Using this theoretical yield and the provided value for actual yield, the percent yield is calculated to be

$$\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100$$

$$\text{percent yield} = \left(\frac{0.392\ \mathrm{g\ Cu}}{0.5072\ \mathrm{g\ Cu}}\right) \times 100$$

$$= 77.3\%$$

Exercise 9.4a

What is the percent yield of a reaction that produces 12.5 g of the gas Freon CF_2Cl_2 from 32.9 g of CCl_4 and excess HF?

$$\mathrm{CCl_4} + 2\mathrm{HF} \longrightarrow \mathrm{CF_2Cl_2} + 2\mathrm{HCl}$$

Check Your Answer[1]

Green Chemistry and Atom Economy

The purposeful design of chemical products and processes that minimize the use of environmentally hazardous substances and the generation of waste is known as *green chemistry*. Green chemistry is a philosophical approach that is being applied to many areas of science and technology, and its practice is summarized by guidelines known as the "Twelve Principles of Green Chemistry (https://www.acs.org/greenchemistry/principles/12-principles-of-green-chemistry.html)". One of the 12 principles is aimed specifically at maximizing the efficiency of processes for synthesizing chemical products. The *atom economy* of a process is a measure of this efficiency, defined as the percentage by mass of the final product of a synthesis relative to the masses of *all* the reactants used:

$$\text{atom economy} = \frac{\text{mass of product}}{\text{mass of reactants}} \times 100\%$$

Though the definition of atom economy at first glance appears very similar to that for percent yield, be aware that this property represents a difference in the *theoretical* efficiencies of *different* chemical processes. The percent yield of a given chemical process, on the other hand, evaluates the efficiency of a process by comparing the yield of product actually obtained to the maximum yield predicted by stoichiometry.

The synthesis of the common nonprescription pain medication, ibuprofen, nicely illustrates the success of a green chemistry approach (Figure 9.4a). First marketed in the early 1960s, ibuprofen was produced using a six-step synthesis that required 514 g of reactants to generate each mole (206 g) of ibuprofen, an atom economy of 40%. In the 1990s, an alternative process was developed by the BHC Company (now BASF Corporation) that requires only three steps and has an atom economy of ~80%, nearly twice that of the original process. The BHC process generates significantly less chemical waste; uses less-hazardous and recyclable materials; and provides significant cost-savings to the manufacturer (and, subsequently, the consumer). In recognition of the positive environmental impact of the BHC process, the company received the Environmental Protection Agency's Greener Synthetic Pathways Award in 1997.

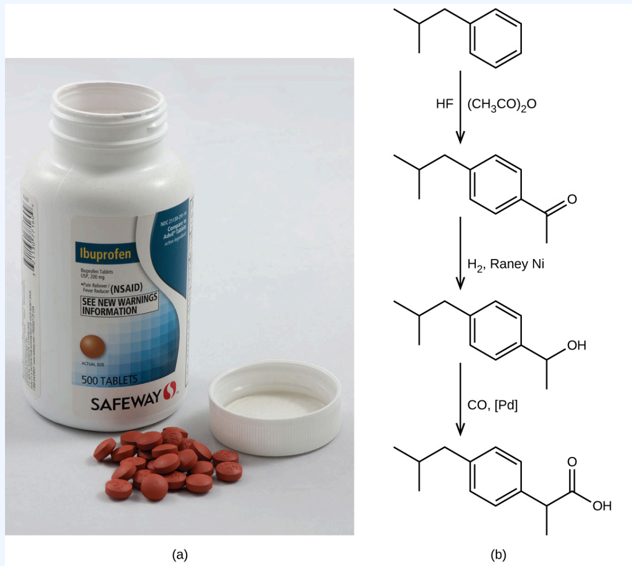

Figure 9.4a (a) Ibuprofen is a popular nonprescription pain medication commonly sold as 200 mg tablets. (b) The BHC process for synthesizing ibuprofen requires only three steps and exhibits an impressive atom economy. (credit a: modification of work by Chiara Coetzee, CC0)

Key Equations

- $\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100$

Attribution & References

Notes

1. 48.3%

CHAPTER 9 - SUMMARY

9.1 Stoichiometric Basics

A balanced chemical equation may be used to describe a reaction's stoichiometry (the relationships between amounts of reactants and products). Coefficients from the equation are used to derive stoichiometric factors that subsequently may be used for computations relating reactant and product masses, molar amounts, and other quantitative properties. Mole quantities of one substance can be related to mass quantities using a balanced chemical equation.

9.2 Mole-Mass and Mass-Mass Calculations

Mole quantities of one substance can be related to mass quantities using a balanced chemical equation. Mass quantities of one substance can be related to mass quantities using a balanced chemical equation. In all cases, quantities of a substance must be converted to moles before the balanced chemical equation can be used to convert to moles of another substance.

9.3 Limiting Reactants

When reactions are carried out using less-than-stoichiometric quantities of reactants, the amount of product generated will be determined by the limiting reactant.

9.4 Reaction Yields

The amount of product generated by a chemical reaction is its actual yield. This yield is often less than the amount of product predicted by the stoichiometry of the balanced chemical equation representing the reaction (its theoretical yield). The extent to which a reaction generates the theoretical amount of product is expressed as its percent yield.

Attribution & References

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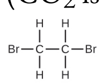
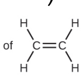
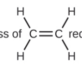
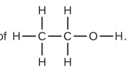
- " (https://opentextbc.ca/introductorychemistry/part/chapter-3-atoms-molecules-and-ions/)Chapter 5: Stoichiometry and The Mole: Mole-Mass and Mass-Mass Calculations" In *Introductory Chemistry: 1st Canadian Edition* by David W. Ball and Jessica A. Key, licensed under CC BY-NC-SA 4.0. / Key takeaways section reused.

CHAPTER 9 - REVIEW

9.1 Stoichiometry Basics; and 9.2 Mole-Mass and Mass-Mass Calculations

1. Write the balanced equation, then outline the steps necessary to determine the information requested in each of the following:
   a. The number of moles and the mass of chlorine, Cl_2, required to react with 10.0 g of sodium metal, Na, to produce sodium chloride, NaCl.
   b. The number of moles and the mass of oxygen formed by the decomposition of 1.252 g of mercury(II) oxide.
   c. The number of moles and the mass of sodium nitrate, $NaNO_3$, required to produce 128 g of oxygen. ($NaNO_2$ is the other product.)
   d. The number of moles and the mass of carbon dioxide formed by the combustion of 20.0 kg of carbon in an excess of oxygen.
   e. The number of moles and the mass of copper(II) carbonate needed to produce 1.500 kg of copper(II) oxide. (CO_2 is the other product.)
   f. The number of moles and the mass of  formed by the reaction of 12.85 g of  with an excess of Br_2.
2. Determine the number of moles and the mass requested for each reaction in Q.1 a) to f) above. **Check Answer:** [1]
3. Write the balanced equation, then outline the steps necessary to determine the information requested in each of the following
   a. The number of moles and the mass of Mg required to react with 5.00 g of HCl and produce $MgCl_2$ and H_2.
   b. The number of moles and the mass of oxygen formed by the decomposition of 1.252 g of silver(I) oxide.
   c. The number of moles and the mass of magnesium carbonate, $MgCO_3$, required to produce 283 g of carbon dioxide. (MgO is the other product.)
   d. The number of moles and the mass of water formed by the combustion of 20.0 kg of acetylene, C_2H_2, in an excess of oxygen.
   e. The number of moles and the mass of barium peroxide, BaO_2, needed to produce 2.500 kg of barium oxide, BaO (O_2 is the other product.)

f. The number of moles and the mass of  required to react with H_2O to produce 9.55 g of .

4. Determine the number of moles and the mass requested for each reaction in Q3. a) to f) above. **Check Answer:** [2]
5. I_2 is produced by the reaction of 0.4235 mol of $CuCl_2$ according to the following equation: $2CuCl_2 + 4KI \longrightarrow 2CuI + 4KCl + I_2$.

   (a) What mass of I_2 is produced?
6. Silver is often extracted from ores such as $K[Ag(CN)_2]$ and then recovered by the reaction $2K[Ag(CN)_2](aq) + Zn(s) \longrightarrow 2Ag(s) + Zn(CN)_2(aq) + 2KCN(aq)$ (a) What mass of $Zn(CN)_2$ is produced? **Check Answer:** [3]
7. What mass of CO_2 is produced by the combustion of 1.00 mol of CH_4? **Check Answer:** [4]

   $$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)$$
8. What mass of H_2O is produced by the combustion of 1.00 mol of CH_4?

   $$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)$$
9. What mass of HgO is required to produce 0.692 mol of O_2? **Check Answer:** [5]

   $$2HgO(s) \rightarrow 2Hg(\ell) + O_2(g)$$
10. What mass of $NaHCO_3$ is needed to produce 2.659 mol of CO_2?

    $$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(\ell) + CO_2(g)$$
11. How many moles of Al can be produced from 10.87 g of Ag? **Check Answer:** [6]

    $$Al(NO_3)_3(s) + 3Ag \rightarrow Al + 3AgNO_3$$
12. How many moles of HCl can be produced from 0.226 g of $SOCl_2$?

    $$SOCl_2(\ell) + H_2O(\ell) \rightarrow SO_2(g) + 2HCl(g)$$
13. How many moles of O_2 are needed to prepare 1.00 g of $Ca(NO_3)_2$? **Check Answer:** [7]

    $$Ca(s) + N_2(g) + 3O_2(g) \rightarrow Ca(NO_3)_2(s)$$
14. How many moles of C_2H_5OH are needed to generate 106.7 g of H_2O?

    $$C_2H_5OH(\ell) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(\ell)$$
15. What mass of O_2 can be generated by the decomposition of 100.0 g of $NaClO_3$? **Check Answer:** [8]

    $$2NaClO_3 \rightarrow 2NaCl(s) + 3O_2(g)$$
16. What mass of Li_2O is needed to react with 1,060 g of CO_2?

$$Li_2O(aq) + CO_2(g) \rightarrow Li_2CO_3(aq)$$

17. What mass of Fe_2O_3 must be reacted to generate 324 g of Al_2O_3? **Check Answer:** [9]
$$Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(s) + Al_2O_3(s)$$

18. What mass of Fe is generated when 100.0 g of Al are reacted?
$$Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(s) + Al_2O_3(s)$$

19. What mass of MnO_2 is produced when 445 g of H_2O are reacted? **Check Answer:** [10]
$$H_2O(\ell) + 2MnO_4^-(aq) + Br^-(aq) \rightarrow BrO_3^-(aq) + 2MnO_2(s) + 2OH^-(aq)$$

20. What mass of $PbSO_4$ is produced when 29.6 g of H_2SO_4 are reacted?
$$Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(\ell)$$

21. If 83.9 g of ZnO are formed, what mass of Mn_2O_3 is formed with it? **Check Answer:** [11]
$$Zn(s) + 2MnO_2(s) \rightarrow ZnO(s) + Mn_2O_3(s)$$

22. If 14.7 g of NO_2 are reacted, what mass of H_2O is reacted with it?
$$3NO_2(g) + H_2O(\ell) \rightarrow 2HNO_3(aq) + NO(g)$$

23. If 88.4 g of CH_2S are reacted, what mass of HF is produced? **Check Answer:** [12]
$$CH_2S + 6F_2 \rightarrow CF_4 + 2HF + SF_6$$

24. If 100.0 g of Cl_2 are needed, what mass of NaOCl must be reacted?
$$NaOCl + HCl \rightarrow NaOH + Cl_2$$

25. What mass of silver oxide, Ag_2O, is required to produce 25.0 g of silver sulfadiazine, $AgC_{10}H_9N_4SO_2$, from the reaction of silver oxide and sulfadiazine?
$$\mathbf{2C_{10}H_{10}N_4SO_2 + Ag_2O \longrightarrow 2AgC_{10}H_9N_4SO_2 + H_2O}$$

26. Carborundum is silicon carbide, SiC, a very hard material used as an abrasive on sandpaper and in other applications. It is prepared by the reaction of pure sand, SiO_2, with carbon at high temperature. Carbon monoxide, CO, is the other product of this reaction. Write the balanced equation for the reaction, and calculate how much SiO_2 is required to produce 3.00 kg of SiC. **Check Answer:** [13]

9.3 Limiting Reactants and 9.4 Reaction Yields

1. Urea, $CO(NH_2)_2$, is manufactured on a large scale for use in producing urea-formaldehyde plastics and as a fertilizer. What is the maximum mass of urea that can be manufactured from the CO_2 produced by combustion of 1.00×10^3kg of carbon followed by the reaction? **Check Answer:** [14]
$$CO_2(g) + 2NH_3(g) \longrightarrow CO(NH_2)_2(s) + H_2O(l)$$

2. In an accident, a solution containing 2.5 kg of nitric acid was spilled. Two kilograms of Na_2CO_3 was quickly spread on the area and CO_2 was released by the reaction. Was sufficient Na_2CO_3 used to neutralize all of the acid?
3. A compact car gets 37.5 miles per gallon on the highway. If gasoline contains 84.2% carbon by mass and has a density of 0.8205 g/mL, determine the mass of carbon dioxide produced during a 500-mile trip (3.785 litres per gallon).
4. What is the limiting reactant in a reaction that produces sodium chloride from 8 g of sodium and 8 g of diatomic chlorine? **Check Answer:** [15]
5. Which of the postulates of Dalton's atomic theory explains why we can calculate a theoretical yield for a chemical reaction?
6. A student isolated 25 g of a compound following a procedure that would theoretically yield 81 g. What was his percent yield? **Check Answer:** [16]
7. A sample of 0.53 g of carbon dioxide was obtained by heating 1.31 g of calcium carbonate. What is the percent yield for this reaction?
   $$\mathrm{CaCO_3}(s) \longrightarrow \mathrm{CaO}(s) + \mathrm{CO_2}(s)$$
8. Freon-12, CCl_2F_2, is prepared from CCl_4 by reaction with HF. The other product of this reaction is HCl. Outline the steps needed to determine the percent yield of a reaction that produces 12.5 g of CCl_2F_2 from 32.9 g of CCl_4. Freon-12 has been banned and is no longer used as a refrigerant because it catalyzes the decomposition of ozone and has a very long lifetime in the atmosphere. Determine the percent yield. **Check Answer:** [17]
9. Citric acid, $C_6H_8O_7$, a component of jams, jellies, and fruity soft drinks, is prepared industrially via fermentation of sucrose by the mold *Aspergillus niger*. The equation representing this reaction is
   $$\mathrm{C_{12}H_{22}O_{11}} \ + \ \mathrm{H_2O} \ + 3\mathrm{O_2} \longrightarrow 2\mathrm{C_6H_8O_7} \ + \ 4\mathrm{H_2O}$$
   What mass of citric acid is produced from exactly 1 metric ton (1.000×10^3 kg) of sucrose if the yield is 92.30%?
10. Toluene, $C_6H_5CH_3$, is oxidized by air under carefully controlled conditions to benzoic acid, $C_6H_5CO_2H$, which is used to prepare the food preservative sodium benzoate, $C_6H_5CO_2Na$. What is the percent yield of a reaction that converts 1.000 kg of toluene to 1.21 kg of benzoic acid? **Check Answer:** [18]
    $$2\mathrm{C_6H_5CH_3} \ + \ 3\mathrm{O_2} \longrightarrow 2\mathrm{C_6H_5CO_2H} \ + \ 2\mathrm{H_2O}$$
11. In a laboratory experiment, the reaction of 3.0 mol of H_2 with 2.0 mol of I_2 produced 1.0 mol of HI. Determine the theoretical yield in grams and the percent yield for this reaction.
12. Outline the steps needed to determine the limiting reactant when 30.0 g of propane, C_3H_8, is burned with 75.0 g of oxygen. Determine the limiting reactant.
13. Outline the steps needed to determine the limiting reactant when 0.50 mol of Cr and 0.75 mol of H_3PO_4 react according to the following chemical equation.
    $$2\mathrm{Cr} \ + \ 2\mathrm{H_3PO_4} \longrightarrow 2\mathrm{CrPO_4} \ + \ 3\mathrm{H_2}$$

Determine the limiting reactant. **Check Answer:** [19]

14. What is the limiting reactant when 1.50 g of lithium and 1.50 g of nitrogen combine to form lithium nitride, a component of advanced batteries, according to the following unbalanced equation?
15. Uranium can be isolated from its ores by dissolving it as $UO_2(NO_3)_2$, then separating it as solid $UO_2(C_2O_4)·3H_2O$. Addition of 0.4031 g of sodium oxalate, $Na_2C_2O_4$, to a solution containing 1.481 g of uranyl nitrate, $UO_2(NO_3)_2$, yields 1.073 g of solid $UO_2(C_2O_4)·3H_2O$.

    $Na_2C_2O_4 + UO_2(NO_3)_2 + 3H_2O \longrightarrow UO_2(C_2O_4){\cdot}3H_2O + 2NaNO_3$

    Determine the limiting reactant and the percent yield of this reaction. **Check Answer:** [20]
16. The phosphorus pentoxide used to produce phosphoric acid for cola soft drinks is prepared by burning phosphorus in oxygen.
    a. What is the limiting reactant when 0.200 mol of P_4 and 0.200 mol of O_2 react according to
       $$P_4 + 5O_2 \longrightarrow P_4O_{10}$$
    b. Calculate the percent yield if 10.0 g of P_4O_{10} is isolated from the reaction.

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- " (https://opentextbc.ca/introductorychemistry/part/chapter-3-atoms-molecules-and-ions/)Chapter 5: Stoichiometry and The Mole: Mole-Mass and Mass-Mass Calculations" In *Introductory Chemistry: 1st Canadian Edition* by David W. Ball and Jessica A. Key, licensed under CC BY-NC-SA 4.0. / Adapted exercises 7-24 for this page.

Notes

1. (a) 0.435 mol Na, 0.217 mol Cl_2, 15.4 g Cl_2; (b) 0.005780 mol HgO, 2.890×10^{-3} mol O_2, 9.248×10^{-2} g O_2; (c) 8.00 mol $NaNO_3$, 6.8×10^2 g $NaNO_3$; (d) 1665 mol CO_2, 73.3 kg CO_2; (e) 18.86 mol CuO, 2.330 kg $CuCO_3$; (f) 0.4580 mol $C_2H_4Br_2$, 86.05 g $C_2H_4Br_2$
2. (a) 0.0686 mol Mg, 1.67 g Mg; (b) 2.701×10^{-3} mol O_2, 0.08644 g O_2; (c) 6.43 mol $MgCO_3$, 542 g $MgCO_3$ (d) 713 mol H_2O, 12.8 kg H_2O; (e) 16.31 mol BaO_2, 2762 g BaO_2; (f) 0.207 mol C_2H_4, 5.81 g C_2H_4
3. (a) 10.41 g $Zn(CN)_2$
4. 44.0 g
5. 3.00×10^2 g

6. 0.0336 mol
7. 0.0183 mol
8. 45.1 g
9. 507 g
10. 4.30×10^3 g
11. 163 g
12. 76.7 g
13. $SiO_2 + 3C \longrightarrow SiC + 2CO$, 4.50 kg SiO_2
14. 5.00×10^3 kg
15. The limiting reactant is Cl_2.
16. Percent yield = 31%
17. g CCl_4 → mol CCl_4 → mol CCl_2F_2 → g CCl_2F_2, percent yield = 48.3%
18. percent yield = 91.3%
19. The conversion needed is mol Cr→ mol H_3PO_4. Then compare the amount of Cr to the amount of acid present. Cr is the limiting reactant.
20. $Na_2C_2O_4$ is the limiting reactant. percent yield = 86.6%