Learning Objectives

As you work through this chapter you will learn how to:

- calculate the number of atoms or molecules in a sample.
- perform calculations involving mole units.
- use the molar mass of a substance in calculations.
- determine the mass percent composition of a compound.
- determine the theoretical yield of a product in a chemical reaction.
- determine the limiting reactant in a chemical reaction.

5.1 A Sure Bet

Long before television, the Internet and computer games a popular activity among children was the game of marbles. In the basic game, players try to knock out small stone or glass marbles (mibs) from a circular playing area using a slightly larger shooter marble (taw). Typically a player carried a set of glass marbles in a cloth pouch and the game was played on just about any smooth hard surface. Marble tournaments are still held today all around the world. Expressions such as “losing your marbles”, “playing for keeps” and “knuckle down” can be traced back to the game of marbles.

With a little bit of information, you can use a collection of marbles to amaze your friends and win a friendly wager. You’ll need a set of identical glass or stone marbles
and an opaque container like a small cardboard box or even a paper bag.¹ Let’s imagine using a group of marbles and a small cardboard box with lid similar to that shown in Figure 5.1. Ask a friend to secretly place as many marbles in the box as he or she wishes and then close the box. The bet is that you will be able to determine how many marbles are in the box without opening it (and in fact without even touching the box of marbles!).

![Figure 5.1 Set of marbles and small cardboard box](image)

In order to make good on your claim, you will need to know in advance the mass of the container and the typical mass of one of the marbles. Then all that you need to do is to weigh the box of marbles prepared by your friend. At home you can conveniently weigh objects such as the box or one marble using an inexpensive digital scale such as that shown in Figure 5.2. The object to be weighed is simply placed on the scale and the mass is displayed in a digital readout in grams. In order to minimize errors, it is necessary to obtain the mass of the objects to three significant figures. If you use marbles this means obtaining masses to the nearest 0.01 g.

¹You can use any set of small objects that are identical in size, such as paper clips, however, it is better if the mass of the object is significant compared to the container you use. Coins could be used, however, the composition of various coins has changed over the years and so has their mass. For example, pennies minted before 1983 have a different composition than that for those minted after 1983 so a mix of pennies would not work.
Let’s assume you weigh several individual marbles and find that the “average” marble in your collection weighs 5.21 g. You also determine that the box weighs 12.84 g. Since 1 marble corresponds to 5.21 g, this provides you with a conversion factor between mass and marbles. Now your friend secretly places several marbles in the box and then places the box of marbles on the scale and it reads 49.14 g. In order to figure out how many marbles are in the box you first calculate the total mass of the marbles in the box. This is determined by subtracting out the mass of the box.

\[
\text{total marble mass} = (\text{mass of box + marbles}) - (\text{mass of box only})
\]

\[
= 49.14 \text{ g} - 12.84 \text{ g} = 36.30 \text{ g}
\]

Now, to get the unknown number of marbles in the box, you use your conversion factor in the following way.

\[
\text{number of marbles} = \frac{\text{total marble mass (g)}}{5.21 \text{ g}} \times \frac{1\text{ marble}}{5.21 \text{ g}}
\]

\[
= 6.97 \text{ marbles}
\]
Since there must be a whole number of marbles in the box, the closest whole number is 7 so you tell your friend there must be 7 marbles in the box (and you win the bet!). This approach to counting objects by weighing can work for any type of object, even tiny invisible ones. This is exactly how chemists determine the number of atoms or molecules in a sample.

5.2 Counting Invisible Objects

If, for example, you want to determine the number of atoms in a sample of gold using the same approach as was used for the marbles, you need to know the mass of a single gold atom. However, remember from Chapter 3 that an individual atom is very small with a mass on the order of $10^{-22}$ g. It is not so easy to weigh individual atoms as was done for the marbles. In the 1800s chemists, such as John Dalton, were able to assign relative masses, called atomic weights, to atoms. They did this by comparing the masses of two elements that react with each other. Dalton chose hydrogen (assumed to have an atomic weight of 1) as the reference point for his scale of atomic weights. Scientists continue to use a scale of relative atomic masses, however, now the mass of atoms and molecules is determined very precisely using a technique called mass spectrometry. In this approach, the mass of an ionized atom or molecule is determined by monitoring its motion in a magnetic field. The present system of atomic weights compares the mass of an atom to $\frac{1}{12}$ the atomic mass of one carbon-12 ($^{12}\text{C}$) atom.

Thus, an atomic weight ($A_r$) is a ratio of masses $\frac{\text{mass of one atom}}{\frac{1}{12}\text{ mass of one }^{12}\text{C atom}}$ and has no units associated with it.\(^2\)


\(^3\)The symbol AW is often used by chemists to designate atomic weight.
Since the mass of individual atoms is a very, very small number when expressed in SI units such as g or kg, it is convenient to express atomic masses using the atomic mass unit (u). One atomic mass unit is defined as exactly 1/12 the mass of a single $^{12}$C atom. Thus, one $^{12}$C atom weighs exactly 12 u. The atomic mass unit has been determined to correspond to a value of $1.6605 \times 10^{-24}$ g. With this unit, the mass of a proton (1.0073 u) and the mass of a neutron (1.0087 u) are just slightly larger than one atomic mass unit and the mass of an electron is 0.00054858 u.

The atomic weight of gold is 197.0. Since this atomic weight represents a ratio of

\[
\frac{\text{mass of one gold atom (in u)}}{\text{mass of one }^{12}\text{C atom (}= 1\text{u})}
\]

gold atom in atomic mass units. This information can now be used to determine the number of gold atoms in a sample of pure gold, as shown in Example 5.1.

---

4The official name is unified atomic mass unit. An equivalent unit is the dalton (Da) which is commonly used in biochemistry; 1 Da = 1 u.

5The atomic mass unit equivalent in grams has been determined to 10 significant figures. Atomic weights of many elements are known to 6 or more significant figures.
CHAPTER 5     HOW CHEMISTS MEASURE AtOMS AND MOLECULES

**Example 5.1**

*Problem*

Determine the number of gold atoms in one ounce (28.35 g) of gold (Au).

*Solution*

What we know: g Au; atomic weight of Au

Desired answer: number of Au atoms

The mass of gold can be converted into number of gold atoms using the atomic weight of gold. Recall that the numerical value of the atomic weight of gold refers to the mass of a single gold atom in atomic mass units. First, convert this value to an equivalent mass in grams.

\[
\frac{197.0 \text{ u Au}}{1 \text{ Au atom}} \times \frac{1.6605 \times 10^{-24} \text{ g Au}}{1 \text{ u Au}} = \frac{3.271 \times 10^{-22} \text{ g Au}}{\text{Au atom}}
\]

Now use this result as the conversion factor between mass of gold and atoms of gold.

\[
28.35 \text{ g Au} \times \frac{1 \text{ Au atom}}{3.271 \times 10^{-22} \text{ g Au}} = 8.667 \times 10^{22} \text{ atoms Au}
\]

Note: As you might have expected, the number of these very small atoms in our sample is enormous.

**Check for Understanding 5.1**

1. The atomic weight of sodium is 22.99. What is the mass, in grams, of a billion sodium atoms?

This approach to relating the number of atoms of an element to the sample mass using the mass of one atom works fine for elements like gold and sodium. For most elements, however, there is an added complication. For instance, if you were to carefully analyze a sample of copper wire using mass spectrometry you would discover that not all
of the copper atoms in the sample have the same mass. One type of copper atom weighs 62.9296 u and another weighs 64.9278 u. Copper, unlike gold, has two stable atomic forms called isotopes. Isotopes are atoms of the same element that have different masses. In order for these atoms to be of the same element they must have the same atomic number (number of protons) and the same number of electrons. Thus, the difference in mass must result from the atoms having different numbers of neutrons (and hence different mass numbers). For copper, the two stable isotopes are $^{63}\text{Cu}$ and $^{65}\text{Cu}$.

So what does the atomic weight of copper refer to, and how can we relate the mass of a copper sample to number of copper atoms as we did for gold?

To answer these questions, let’s return to the marble example briefly. Recall that we were using “identical” marbles, but suppose your friend had placed some of the larger shooter marbles in the box with the smaller ones. Now your average marble mass does not represent the mix of marbles in the box and your calculated number of marbles will be in error. What you need is the average mass of a marble in this mix of marbles. Let’s assume that your friend used a 50-50 mix; half are identical small marbles and half are identical large marbles. The average mass of a marble in this mix would be the average of the two marble masses. This can be represented mathematically as:

\[
\text{average marble mass (50:50 mix)} = \frac{1}{2} (\text{smaller marble mass} + \text{larger marble mass})
\]

However, your friend may have used only one larger shooter and the rest small marbles. To determine the average marble mass for any combination of large and small marbles, notice that the $\frac{1}{2}$ factor in the equation above represents the fraction of that particular type of marble. Thus, to get the average marble mass for any mix you can use the following relationship.

\[
\text{average marble mass} = (\text{fraction of small marbles})(\text{smaller marble mass}) + (\text{fraction of large marbles})(\text{larger marble mass})
\]
The average marble mass calculated in this way is known as a **weighted average**. A weighted average can be calculated for as many different types of marbles as you wish. All that you need to know is the mass of each marble type and the fraction of the total for this marble type.

For elements that have more than one stable isotope\(^6\), a weighted average is used for the mass of one atom of that element. In the case of copper, 69.15% of the atoms in a typical copper sample are \(^{63}\text{Cu}\) and 30.85% are \(^{65}\text{Cu}\). Thus, the weighted average mass of a copper atom is calculated by:

\[
\text{weighted average mass of Cu atoms} = (\text{fractional isotope abundance})\left(^{63}\text{Cu mass}\right) + (\text{fractional isotope abundance})\left(^{65}\text{Cu mass}\right)
\]

\[
= (0.6915)(62.9296 \text{ u}) + (0.3085)(64.9278 \text{ u}) = 63.55 \text{ u}
\]

Therefore, the atomic weight of copper is 63.55. This value can be used to determine the number of copper atoms in a sample of copper as was done for gold.

Take note that no atom of copper weighs 63.55 u. This weighted average mass is useful only if the fractional abundance of the copper isotopes is the same for all samples of copper. This same caution applies to all elements that have more than one stable isotope. Scientists have determined that for some elements there is a significant variation in the isotopic composition from one type of sample to another. Starting in 2011, the atomic weights of 10 elements (H, Li, B, C, N, O, Si, S, Cl and Tl) are expressed as intervals that reflect this variation in isotopic abundance. For most instructional purposes, including this course, the conventional atomic weights for these elements are sufficient. The periodic table shown on the next page includes the atomic weights to be used in this course. Generally you will be provided a periodic table with such numerical information for use on quizzes and exams. It is very important that you understand and can apply all of the information available from this periodic table.

\(^6\)Only 26 elements have a single stable isotope.
Figure 5.3 Periodic Table of the elements showing atomic weights

*Elements which have no known stable isotopes (that is, all atoms of the element are radioactive) and are not found in nature have the atomic weight of the most stable isotope in brackets.
5.3 The Mole

Since a very large number of atoms (and molecules) are associated with typical samples of material, scientists have created a unit that allows such quantities to be more conveniently expressed. This unit is the mole (mol). The mole is a counting unit like the dozen and is linked to the reference point ($^{12}$C) of the atomic weight system. One mole is defined as the number of atoms of $^{12}$C in exactly 12 grams of $^{12}$C (Note that 12 grams of carbon corresponds to a 60-carat diamond!). The Avogadro constant ($N_A$), often referred to as Avogadro’s number, expresses the number of elementary entities (atoms, molecules, etc.) there are per mole of a substance. It has been experimentally determined to have a value (to 4 significant figures) of $6.022 \times 10^{23}$ mol$^{-1}$ (Note that mol$^{-1}$ means the same as 1/mol.). Thus, one mole of anything is the amount of that substance that contains $6.022 \times 10^{23}$ items. The value of the Avogadro constant is so large it is difficult to comprehend just how many items are associated with a mole of something. For example, if we take a mole of pennies and distribute these equally to the approximately 7 billion people on earth, then each individual would have nearly a trillion dollars!

Consequently, the mole is a good unit for counting very small objects such as atoms and molecules. Since the mole is so commonplace in chemistry you should memorize the value of the Avogadro constant.

$$N_A = 6.022 \times 10^{23} \text{ mol}^{-1}$$

Since a mole is defined in terms of a number of grams numerically equal to the atomic mass (in u) of $^{12}$C, a similar connection can be made for the other elements. The number of grams numerically equal to the weighted average atomic mass (in u) of an element will contain one mole of atoms of that element. For example, 1 mole of silver weighs 107.9 g and 1 mole of sulfur weighs 32.07 g. We can determine this by noting the atomic weights of these elements listed on the periodic table in Figure 5.3.

The mass of a substance that corresponds to one mole of that material is called the molar mass (M) of that substance. In chemistry molar masses typically are expressed in units of g/mol. This quantity serves as the conversion factor between the mass of a
substance and the moles of that substance. You can use it to convert from either mass to moles or from moles to mass.\(^8\) The Avogadro constant \((N_A)\) is used to convert from moles to a specific number of fundamental units (atoms, molecules, ions) present, or to convert from a number of particles to the corresponding number of moles.

Molar mass will be one of the most frequently used conversion factors in this course.

The steps of the problem-solving strategy described in Section 2.6 for problems involving calculations are listed below. This approach is used to set up and solve the many problems involving calculations in this text. You should review these steps before studying subsequent Examples and before starting work on the Check for Understanding problems and the Exercises in this chapter.

<table>
<thead>
<tr>
<th>Strategy for Solving Calculation Problems</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Read the entire problem very carefully, making note of key terms and units.</td>
</tr>
<tr>
<td>2. Identify the question to be answered and make special note of the units for the answer.</td>
</tr>
<tr>
<td>3. Look at all the information that is given, including units for numerical values.</td>
</tr>
<tr>
<td>4. Think about relationships that connect the given information to the question and its plausible answer and create your solution map, identifying each step.</td>
</tr>
<tr>
<td>5. Identify all the conversion factors needed for your solution map.</td>
</tr>
<tr>
<td>6. Confirm that the relevant units cancel to give you the desired units for the answer.</td>
</tr>
<tr>
<td>7. Do the calculations carefully. Think about whether or not your answer makes sense.</td>
</tr>
</tbody>
</table>

\(^8\)The symbol \(\leftrightarrow\) means a process that can go in either direction.
A key to this strategy is the creation of a solution map that outlines how the starting units are converted to the units of the desired answer. For example, if you are given information about the mass (in g) of substance A and asked for the number of moles of substance A, then the solution map would look like: \( g \text{ A} \rightarrow \text{mol A} \). Since molar mass has units of g/mol, this is the logical choice for the conversion factor needed in this calculation. However, remember that conversion factors can always be expressed in two forms, one the inverse of the other. Thus, molar mass can be used as \( \frac{g}{\text{mol}} \) or as \( \frac{\text{mol}}{g} \). Which form you use depends on the unit conversion needed. In this case, we need a conversion factor with units of \( \frac{\text{mol}}{g} \) which will cancel g and introduce \( \text{mol} \) in the numerator, so the unit conversion looks like:

\[
g \times \frac{\text{mol}}{g} = \text{mol}
\]

**Check for Understanding 5.2**

1. Write the solution map for the conversion of grams of sodium to atoms of sodium. Indicate the numerical ratio that is the conversion factor in each step.

The use of the mole and molar mass in calculations involving elements is illustrated in Examples 5.2 and 5.3.
Example 5.2

Problem

How many iron (Fe) atoms are present in 0.79 mmol of iron?

Solution

What we know: mmol Fe
Desired answer: number of Fe atoms

The solution map for this calculation is:

\[
\text{mmol Fe} \rightarrow \text{mol Fe} \rightarrow \text{atoms Fe}
\]

The conversion factor needed in the first step is the relationship between \( \text{mmol} \) and \( \text{mol} \). Since \( 1 \text{ mmol} = 10^{-3} \text{ mol} \), use \( \frac{10^{-3} \text{ mol Fe}}{1\text{ mmol Fe}} \) in order to convert units properly.

The conversion factor needed in the second step is the Avogadro constant. It is used in the form \( \frac{6.022 \times 10^{23} \text{ atoms Fe}}{1\text{ mol Fe}} \) in order to convert units properly.

Putting these together yields:

\[
0.79 \text{ mmol Fe} \times \frac{10^{-3} \text{ mol Fe}}{1\text{ mmol Fe}} \times \frac{6.022 \times 10^{23} \text{ atoms Fe}}{1\text{ mol Fe}} = 4.8 \times 10^{20} \text{ atoms Fe}
\]

Note that 2 significant figures are appropriate for the answer.
Example 5.3

Problem

What is the mass (in kg) of 18.4 mol of mercury (Hg)?

Solution

What we know: mol Hg

 Desired answer: kg Hg

The solution map for this calculation is

\[ \text{mol Hg} \rightarrow \text{g Hg} \rightarrow \text{kg Hg} \]

The conversion factor needed in the first step is the molar mass of mercury. The numerical value for the molar mass is obtained from the periodic table and is expressed in the form \( \frac{200.6 \text{ g Hg}}{1 \text{ mol Hg}} \) in order to convert units properly.

The conversion factor needed in the second step is the relationship between kg and g. Since \( 1 \text{ kg} = 10^3 \text{ g} \), use \( \frac{1 \text{ kg Hg}}{10^3 \text{ g Hg}} \) in order to convert units properly.

Putting these together yields:

\[ 18.4 \text{ mol Hg} \times \frac{200.6 \text{ g Hg}}{1 \text{ mol Hg}} \times \frac{1 \text{ kg Hg}}{10^3 \text{ g Hg}} = 3.69 \text{ kg Hg} \]

Note that 3 significant figures are appropriate for the answer.

Check for Understanding 5.3

1. What is the mass, in grams, of one cobalt atom?

2. How many atoms of aluminum are present in a piece of aluminum foil that weighs 6.27 g?
5.4 Molar Mass of a Compound

The molar mass of a compound is the mass in grams of one mole of formula units of the compound. The molar mass of a compound is obtained by summing the molar masses of all of the atoms represented in the chemical formula of the compound. If there is more than one atom of an element in the chemical formula, its molar mass is added as many times as the element appears in the formula. The calculation of the molar mass of a compound is illustrated in Examples 5.4 and 5.5.

**Example 5.4**

*Problem*

What is the molar mass of water?

*Solution*

What we know: water; atomic weights of elements

Desired answer: g water/mol

The chemical formula for water is H₂O. The formula indicates that in 1 mole of water molecules there are 2 moles of hydrogen atoms and 1 mole of oxygen atoms. From the periodic table we see that 1 mole of hydrogen atoms weighs 1.008 g and that 1 mole of oxygen atoms weighs 16.00 g. Thus, one mole of water will weigh:

\[
(2 \text{ mol H} \times \frac{1.008 \text{ g}}{1 \text{ mol H}}) + (1 \text{ mol O} \times \frac{16.00 \text{ g}}{1 \text{ mol O}}) = 18.02 \text{ g}
\]

The molar mass of water is 18.02 g/mol.
Example 5.5

Problem

What is the molar mass of calcium nitrate?

Solution

What we know: calcium nitrate; atomic weights of elements

Desired answer: g calcium nitrate/mol

The chemical formula for calcium nitrate is Ca(NO₃)₂. The formula indicates that in 1 mole of calcium nitrate there are 1 mole of calcium, 2 moles of nitrogen and 6 moles of oxygen. From the periodic table we see that 1 mole of calcium atoms weighs 40.08 g, 1 mole of nitrogen atoms weighs 14.01 g and 1 mole of oxygen atoms weighs 16.00 g. Thus, one mole of Ca(NO₃)₂ will weigh:

\[
\begin{align*}
1 \text{ Ca} & \quad 1\text{ molCa} \times \frac{40.08\text{ g}}{1\text{ molCa}} = 40.08\text{ g} \\
2 \text{ N} & \quad 2\text{ molN} \times \frac{14.01\text{ g}}{1\text{ molN}} = 28.02\text{ g} \\
6 \text{ O} & \quad 6\text{ molO} \times \frac{16.00\text{ g}}{1\text{ molO}} = 96.00\text{ g} \\
\text{Total} & \quad 164.10\text{ g} 
\end{align*}
\]

The molar mass of calcium nitrate is 164.10 g/mol.

Check for Understanding 5.4

Solution

1. Calculate the molar mass of sodium phosphate (Na₃PO₄).
Analogous to the situation for chemical elements, the molar mass of a compound connects the macroscopic world to the invisible microscopic one by serving as the conversion factor between a readily measurable mass of a compound and the number of moles of that compound. You can use it to convert from mass of a compound to moles of a compound, or from moles of a compound to the corresponding mass of the compound. The Avogadro constant ($N_A$) is used to convert from moles to a specific number of fundamental units (molecules, formula units) present, or to convert from a number of molecules or formula units to the corresponding number of moles of a compound.

Examples 5.6, 5.7 and 5.8 illustrate how to use the molar mass of a compound in calculations.
Example 5.6

Problem

What is the mass in grams of 0.295 mol MgCl₂?

Solution

What we know: \( \text{mol MgCl}_2 \)

Desired answer: \( \text{g MgCl}_2 \)

The solution map for this calculation is:

\[
\text{mol MgCl}_2 \rightarrow \text{g MgCl}_2
\]

The conversion factor needed is the molar mass of MgCl₂. From the periodic table we can get the molar masses of magnesium and chlorine and add them as follows.

\[
24.31 \text{ g Mg} + 2(35.45 \text{ g}) \text{ Cl} = 95.21 \text{ g}
\]

Thus, the molar mass of MgCl₂ is 95.21 g/mol. It is used in the form \( \frac{95.21 \text{g MgCl}_2}{1 \text{mol MgCl}_2} \) in order to convert units properly.

Applying this yields:

\[
0.295 \text{ mol MgCl}_2 \times \frac{95.21 \text{g MgCl}_2}{1 \text{mol MgCl}_2} = 28.1 \text{ g MgCl}_2
\]
Example 5.7

Problem

How many moles of CH₃OH molecules are there in 0.19 g of CH₃OH?

Solution

What we know: g CH₃OH

Desired answer: mol CH₃OH

The solution map for this calculation is:

\[ \text{g CH}_3\text{OH} \rightarrow \text{mol CH}_3\text{OH} \]

The conversion factor needed is the molar mass of CH₃OH. From the periodic table we can get the molar masses of carbon, hydrogen and oxygen and add them as follows. 12.01 g C + 4(1.008 g) H + 16.00 g O = 32.04 g. Thus, the molar mass of CH₃OH is 32.04 g/mol. It is used in the form \( \frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g CH}_3\text{OH}} \) in order to convert units properly.

Applying this yields:

\[ 0.19 \text{ g CH}_3\text{OH} \times \frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g CH}_3\text{OH}} = 0.0059 \text{ mol CH}_3\text{OH} \]
Example 5.8

Problem

How many molecules of CO$_2$ are there in 112 g of CO$_2$?

Solution

What we know: g CO$_2$

Desired answer: molecules CO$_2$

The solution map for this calculation is:

\[ g \text{ CO}_2 \rightarrow \text{ mol CO}_2 \rightarrow \text{ molecules CO}_2 \]

The conversion factor needed in the first step is the molar mass of CO$_2$. From the periodic table we can get the molar masses of carbon and oxygen and add them as follows. 12.01 g C + 2(16.00 g) O = 44.01 g. Thus, the molar mass of CO$_2$ is 44.01 g/mol. It is used in the form \[ \frac{1\text{molCO}_2}{44.01\text{g CO}_2} \] in order to convert units properly.

The conversion factor needed in the second step is the Avogadro constant. It is used in the form \[ \frac{6.022 \times 10^{23} \text{ molecules CO}_2}{1\text{mol CO}_2} \] in order to convert units properly.

Putting these together yields:

\[
112 \text{ g CO}_2 \times \frac{1\text{molCO}_2}{44.01\text{g CO}_2} \times \frac{6.022 \times 10^{23} \text{ molecules CO}_2}{1\text{mol CO}_2} = 1.53 \times 10^{24} \text{ molecules CO}_2
\]
5.5 Mass Percent Composition of Compounds

Since compounds always have the same specific composition, it is possible to establish quantitative relationships that are useful conversion factors. For example, since water has the composition H₂O, we note there are 2 atoms of H for each atom of O. Such a relationship can be expressed as an atom ratio, and can be “scaled up” by multiplying both numerator and denominator by the Avogadro constant.

For water there are: \[ \frac{2 \text{ atoms H}}{1 \text{ atom O}} \quad \text{or} \quad \frac{2 \text{ mol H}}{1 \text{ mol O}} \]

For water it is also true that: \[ \frac{2 \text{ atoms H}}{1 \text{ molecule H}_2\text{O}} \quad \text{or} \quad \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \]

Example 5.9 illustrates how such relationships can be applied to useful conversions.

Check for Understanding 5.5

1. Calculate the number of moles in a sample of sodium bicarbonate (NaHCO₃) that weighs 6.22 g.

2. How many molecules are present in 1.05 g of iodine (I₂)?

3. What is the mass in grams of 25 molecules of CCl₂F₂?

Check for Understanding 5.6

1. What is the numerical ratio between the number of hydrogen atoms and oxygen atoms in (NH₄)₃PO₄?

2. What is the numerical ratio between the number of moles of sulfur and moles of Al₂(SO₄)₃?
**Example 5.9**

*Problem*

How many moles of hydrogen atoms are in 3.6 mol of ammonium carbonate?

*Solution*

What we know: mol ammonium carbonate

Desired answer: mol H

The solution map for this calculation is:

\[ \text{mol ammonium carbonate} \rightarrow \text{mol H} \]

As in this case, you will often need to provide a correct chemical formula so continue to review chemical names and formulas. The chemical formula for ammonium carbonate is \((NH_4)_2CO_3\). The formula indicates that in 1 mole of ammonium carbonate there are 8 moles of hydrogen atoms. This relationship can be applied as follows.

\[
3.6 \text{ mol (NH}_4\text{)}_2\text{CO}_3 \times \frac{8 \text{ mol H}}{1 \text{ mol (NH}_4\text{)}_2\text{CO}_3} = 29 \text{ mol H}
\]

**Check for Understanding 5.7**

1. Calculate the number of moles of barium nitrate that contain 0.81 mol of oxygen atoms.

2. How many carbon atoms are present in 16.0 g C\(_3\)H\(_8\)?

3. How many grams of silver are present in 25.0 g AgBr?
Another very useful piece of information that is associated with the definite composition of compounds is **percent** (%), or parts per hundred. A number percent refers to the number of items of a specified type in a group of 100 items total. For example, if 12% of the students in a class are left-handed, this means 12/100, or 0.12, of the class has this characteristic. It is important to note that the percent sign (%) relates to implicit units, so if this quantity is to be used as a conversion factor it is important to indicate the specific items and the total items referred to. If 12% of the class is left-handed, then there are 12 left-handed students/100 students in the class. This means that a number percent is calculated by:

\[
\text{number percent} = \frac{\text{no. items of interest}}{\text{total no. of items}} \times 100
\]

Example 5.10 illustrates the use of number percent.

In chemistry, **mass percent** is often a more useful quantity. A mass percent is calculated by:

\[
\text{mass percent} = \frac{\text{mass of component of interest}}{\text{total mass}} \times 100
\]

Note that the mass of the component and the total mass must be expressed in the same units.
The specific composition of a compound is frequently expressed in terms of the mass percent of each constituent element. For example, NaCl is 39.34% sodium (Na) and 60.66% chlorine (Cl) by mass. Notice that the mass percent values total 100% since these two elements are the only ones present. Also, according to the law of constant composition, these mass percent values apply to any sample of sodium chloride. The calculation of the mass percent composition of a compound is illustrated in Example 5.11.

Example 5.10

Problem

If 68% of a class of 93 students are biology majors and 57% of the biology majors are female, how many students in the class are female biology majors?

Solution

What we know: students in class; % biology majors; % female biology majors

Desired answer: number of female biology majors

The solution map for this problem is

students in class \(\rightarrow\) biology majors \(\rightarrow\) female biology majors

The conversion factor needed in the first step is \(\frac{68 \text{ biol. majors}}{100 \text{ students in class}}\).

The conversion factor needed in the second step is \(\frac{57 \text{ female biol. majors}}{100 \text{ biol. majors}}\).

Putting these together yields:

\[
93 \text{ students in class} \times \frac{68 \text{ biol. majors}}{100 \text{ students in class}} \times \frac{57 \text{ female biol. majors}}{100 \text{ biol. majors}} = 36 \text{ female biol. majors}
\]
In summary, you can calculate the mass percent composition of a compound by comparing the total mass of each element present in one mole of compound to the molar mass of the compound. If the formula of the compound is not known, then the mass percent composition can be obtained by analyzing a specific mass of material and measuring the mass of each element present.
5.6 Interpreting Chemical Reactions Quantitatively

We learned that in a chemical reaction atoms are neither created not destroyed, so a balanced chemical equation contains information that you can use to establish quantitative relationships between reactants and products. **Stoichiometry** (pronounced *stoy-kee-ahm-eh-tree*) is the area of chemistry that deals with the quantitative relationships between reactants and products in a chemical reaction. Consider the balanced equation for one of the single displacement reactions described in Chapter 3 in which zinc metal reacts with hydrochloric acid to produce zinc chloride and hydrogen gas.

\[
\text{Zn(s)} + 2\text{HCl(aq)} \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)
\]

At the most fundamental level this equation suggests that 1 atom of zinc reacts with 2 molecules of HCl to produce 1 formula unit of ZnCl\(_2\) and 1 molecule of H\(_2\).

\[
\begin{array}{cccc}
\text{1 atom} & \text{2 molecules} & \text{1 formula unit} & \text{1 molecule}
\end{array}
\]

When such a reaction is carried out in the laboratory, many, many atoms of zinc react with many, many molecules of HCl, however, they always react in a 1:2 ratio.

Check for Understanding 5.8

1. What is the mass percent of carbon in C\(_6\)H\(_{12}\)O\(_6\)?
The balanced equation can be “scaled up” by multiplying each reactant amount and each product amount by the Avogadro constant to yield the following quantitative relationship.

\[
\text{Zn(s)} + 2\text{HCl(aq)} \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g})
\]

\[
1\text{ mol} \quad 2\text{ mol} \quad 1\text{ mol} \quad 1\text{ mol}
\]

These mole quantities of reactants and products can be used to create mole ratios. A **mole ratio** is a ratio of moles of any two substances involved in a chemical reaction. The coefficients in the balanced equation for the reaction are used to establish the numerical part of this ratio. For the above equation, there are 12 possible mole ratios. Look carefully at the equation above to verify each of the 12 ratios given below.

<table>
<thead>
<tr>
<th>1 mol Zn</th>
<th>1 mol Zn</th>
<th>1 mol Zn</th>
<th>2 mol HCl</th>
<th>2 mol HCl</th>
<th>1 mol ZnCl₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>2 mol HCl</td>
<td>1 mol ZnCl₂</td>
<td>1 mol H₂</td>
<td>1 mol ZnCl₂</td>
<td>1 mol H₂</td>
<td>1 mol H₂</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>2 mol HCl</th>
<th>1 mol ZnCl₂</th>
<th>1 mol H₂</th>
<th>1 mol ZnCl₂</th>
<th>1 mol H₂</th>
<th>1 mol H₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mol Zn</td>
<td>1 mol Zn</td>
<td>1 mol Zn</td>
<td>2 mol HCl</td>
<td>2 mol HCl</td>
<td>1 mol ZnCl₂</td>
</tr>
</tbody>
</table>

These ratios can be used to convert between moles of one substance and moles of another substance involved in this reaction. Such **mole-to-mole** calculations are illustrated in Examples 5.12 and 5.13.
Example 5.12

Problem

How many moles of Fe$_2$O$_3$ will be produced from the reaction of 1.00 mol FeS according to the equation below?

$$4\text{FeS(s)} + 7\text{O}_2(\text{g}) \rightarrow 2\text{Fe}_2\text{O}_3(\text{s}) + 4\text{SO}_2(\text{g})$$

Solution

What we know: mol FeS; balanced equation relating FeS and Fe$_2$O$_3$

Desired answer: mol Fe$_2$O$_3$

The solution map for this problem is:

$$\text{mol FeS} \rightarrow \text{mol Fe}_2\text{O}_3$$

The conversion factor needed is the mole ratio for these two substances from the balanced equation in the form \(\frac{2 \text{mol Fe}_2\text{O}_3}{4 \text{mol FeS}}\) in order to convert units properly.

Applying this yields:

$$1.00 \text{ mol FeS} \times \frac{2 \text{mol Fe}_2\text{O}_3}{4 \text{mol FeS}} = 0.500 \text{ mol Fe}_2\text{O}_3$$

Note that the numerical part of a mole ratio involves exact numbers with an unlimited number of significant figures.
Example 5.13

Problem

How many moles of \( O_2 \) are needed to react with 6.50 mol FeS according to the equation below?

\[
4 \text{FeS(s)} + 7 \text{O}_2(\text{g}) \rightarrow 2 \text{Fe}_2\text{O}_3(\text{s}) + 4 \text{SO}_2(\text{g})
\]

Solution

What we know: \( \text{mol FeS} \); balanced equation relating FeS and \( O_2 \)

Desired answer: \( \text{mol } O_2 \)

The solution map for this problem is:

\[
\text{mol FeS} \quad \rightarrow \quad \text{mol } O_2
\]

The conversion factor needed is the mole ratio for these two substances from the balanced equation in the form \( \frac{7 \text{mol } O_2}{4 \text{mol FeS}} \) in order to convert units properly.

Applying this yields:

\[
6.50 \text{ mol FeS} \times \frac{7 \text{ mol } O_2}{4 \text{ mol FeS}} = 11.4 \text{ mol } O_2
\]
In the last few sections of the text you have seen how to form mathematical relationships that allow you to make the following calculations.

<table>
<thead>
<tr>
<th>Calculation</th>
<th>Conversion factor</th>
</tr>
</thead>
<tbody>
<tr>
<td>mol substance $\leftrightarrow$ g substance</td>
<td>molar mass</td>
</tr>
<tr>
<td>mol substance $\leftrightarrow$ number of atoms or molecules</td>
<td>Avogadro constant</td>
</tr>
<tr>
<td>mol compound $\leftrightarrow$ mol constituent element</td>
<td>atom ratio based on formula of compound</td>
</tr>
<tr>
<td>g constituent element $\leftrightarrow$ g compound</td>
<td>mass percent</td>
</tr>
<tr>
<td>mol reactant or product $\leftrightarrow$ mol of another reactant or product</td>
<td>mole ratio from balanced equation</td>
</tr>
</tbody>
</table>

These relationships can be used to form the overall solution map for stoichiometric calculations shown in Figure 5.4. This map applies to the general reaction:

reactant A $\rightarrow$ product B

![Figure 5.4: Solution map for stoichiometric calculations](image)

This map indicates the number of steps and the conversion factors needed for performing various calculations that you will need to make in this course. For example, to convert from molecules of reactant A to mass of product B three steps are required.
The conversion factors are $N_A$, mol B/mol A and the molar mass of B. It is very important that you understand the role of each conversion factor and how to obtain it, and that you practice these calculations by doing many different types of problems. Examples 5.14, 5.15 and 5.16 illustrate applications of this solution map.

**Example 5.14**

**Problem**

How many grams of ammonia (NH$_3$) can be produced from 0.512 mol of hydrogen gas reacting with nitrogen gas according to the equation below?

$$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$$

**Solution**

What we know: mol H$_2$; balanced equation relating NH$_3$ and H$_2$

Desired answer: g NH$_3$

The solution map for this problem is:

$$\text{mol H}_2 \rightarrow \text{mol NH}_3 \rightarrow \text{g NH}_3$$

The conversion factor needed in the first step is the mole ratio for these two substances from the balanced equation in the form $\frac{2\text{ mol NH}_3}{3\text{ mol H}_2}$.

The conversion factor needed in the second step is the molar mass of NH$_3$. From the periodic table we can get the molar masses of nitrogen and hydrogen and add them as follows. 14.01 g N + 3(1.008 g) H = 17.03 g. Thus, the molar mass of NH$_3$ is 17.03 g/mol.

Putting these together yields:

$$0.512 \text{ mol H}_2 \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} \times \frac{17.03 \text{ g NH}_3}{1 \text{ mol NH}_3} = 5.81 \text{ g NH}_3$$
Example 5.15

Problem

How many moles of oxygen gas are produced when 40.6 g of potassium chlorate decomposes according to the equation below?

\[
2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g)
\]

Solution

What we know: g KClO₃; balanced equation relating KClO₃ and O₂

Desired answer: mol O₂

The solution map for this problem is:

\[
g \text{ KClO}_3 \rightarrow \text{ mol KClO}_3 \rightarrow \text{ mol O}_2
\]

The conversion factor needed in the first step is the molar mass of KClO₃. From the periodic table we can get the molar masses of potassium, chlorine and oxygen and add them as follows. 39.10 g K + (35.45 g) Cl + 3(16.00 g) O = 122.55 g. Thus, the molar mass of KClO₃ is 122.55 g/mol. It is used in the form \( \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \) in order to cancel units properly.

The conversion factor needed in the second step is the mole ratio for these two substances from the balanced equation in the form \( \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} \).

Putting these together yields:

\[
40.6 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} = 0.497 \text{ mol O}_2
\]
Example 5.16

Problem

What mass (in grams) of carbon dioxide is produced by the complete combustion of 95.0 grams of propane (C\textsubscript{3}H\textsubscript{8})? The balanced equation for this combustion reaction is shown below.

\[ \text{C}_3\text{H}_8(\text{l}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l}) \]

Solution

What we know: g C\textsubscript{3}H\textsubscript{8}; balanced equation relating C\textsubscript{3}H\textsubscript{8} and CO\textsubscript{2}

Desired answer: g CO\textsubscript{2}

The solution map for this problem is:

\[ \text{g C}_3\text{H}_8 \rightarrow \text{mol C}_3\text{H}_8 \rightarrow \text{mol CO}_2 \rightarrow \text{g CO}_2 \]

The conversion factor needed in the first step is the molar mass of C\textsubscript{3}H\textsubscript{8}. From the periodic table we can get the molar masses of carbon and hydrogen and add them as follows. \( 3(12.01 \text{ g}) C + 8(1.008 \text{ g}) H = 44.09 \text{ g} \). Thus, the molar mass of C\textsubscript{3}H\textsubscript{8} is 44.09 g/mol. It is used in the form \( \frac{1 \text{ mol C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} \) in order to cancel units properly.

The conversion factor needed in the second step is the mole ratio for these two substances from the balanced equation in the form \( \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} \).

The conversion factor needed in the last step is the molar mass of CO\textsubscript{2}. From the periodic table we can get the molar masses of carbon and oxygen and add them as follows. \( 12.01 \text{ g} C + 2(16.00 \text{ g}) O = 44.01 \text{ g} \). Thus, the molar mass of CO\textsubscript{2} is 44.01 g/mol.

Putting these together yields:

\[
95.0 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} \times \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 284 \text{ g CO}_2
\]
CHAPTER 5   HOW CHEMISTS MEASURE ATOMS AND MOLECULES

5.7 Theoretical Yield and Limiting Reactant

If you have ever made a batch of cookies from a detailed recipe you have probably found that the number of cookies you ultimately end up with is very different than what the recipe indicates. This is a result of a number of factors. Cookie size varies, cookie dough is lost on bowls and spoons, some cookies may be discarded because they cooked for too long, and inevitably you have been sampling your delicious product hot from the oven. A similar situation is found for chemical reactions. The balanced equation indicates exactly how much product will form from a specific amount of reactants. The maximum amount of product that can be obtained from a given amount of reactants is called the **theoretical yield** of the reaction. Often, however, side reactions occur that consume reactants and products or you may not be able to fully recover the product that forms so the amount of product that you obtain, called the **actual yield**, is less than expected. An important area of chemistry research involves working to find reactions and the necessary conditions for which the actual yield is very close to the theoretical yield. This is especially important for industrial processes that involve very

---

**Check for Understanding 5.9**

1. Based on the reaction between iron(III) oxide and carbon, how many moles of carbon are needed to produce 1.0 kmol of iron?

\[
\text{Fe}_2\text{O}_3(s) + 3\text{C(s)} \rightarrow 2\text{Fe(s)} + 3\text{CO(g)}
\]

2. What mass of NaHCO₃ will react with 0.00257 mol HCl?

\[
\text{NaHCO}_3(s) + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{CO}_3(aq)
\]

3. How many grams of sodium metal are needed to produce 125 g hydrogen gas by the reaction below?

\[
2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + \text{H}_2(g)
\]
large scale reactions, or processes such as the production of pharmaceuticals that involve many steps.

Often, chemical processes are carried out under conditions in which one reactant is present in excess. This means that you will completely consume the other reactant(s) and will have some of the reactant in excess left over. When one or more of the reactants in a chemical reaction are completely consumed the reaction ceases and no additional product can be formed. A reactant that is completely consumed is called a **limiting reactant** because it limits the amount of product that can form. The maximum amount of product that can form from the limiting reactant is the theoretical yield of a reaction.

To illustrate the idea of a limiting reactant and theoretical yield consider the following situation. Suppose you and your friends wish to put together fruit baskets for gifts. You want each basket to contain 1 grapefruit, 2 apples and 3 bananas. If you collect 8 grapefruit, 14 apples and 27 bananas, how many baskets can you prepare?
At first glance you might imagine that the fruit present in the smallest amount (the grapefruit) would be used up first. This would be the case if you used the same number of each item to make your basket, however, this is not the situation here. In order to determine the maximum number of baskets you can prepare (the theoretical yield), determine how many baskets you could make from each type of fruit.

The yield from the grapefruit is: \[ \frac{8 \text{ grapefruit}}{1 \text{ basket}} \times \frac{1 \text{ basket}}{1 \text{ grapefruit}} = 8 \text{ baskets} \]

The yield from the apples is: \[ \frac{14 \text{ apples}}{1 \text{ basket}} \times \frac{1 \text{ basket}}{2 \text{ apples}} = 7 \text{ baskets} \]

The yield from the bananas is: \[ \frac{27 \text{ bananas}}{1 \text{ basket}} \times \frac{1 \text{ basket}}{3 \text{ bananas}} = 9 \text{ baskets} \]

Once you have made 7 fruit baskets all of the apples will be gone and you can make no more. Thus, the apples are the limiting starting material and the maximum amount of product is 7 baskets. After you have used all of the apples there will be some grapefruit and bananas remaining. To determine the excess of each of these, calculate how many were used to make the 7 baskets and subtract this amount from the starting quantity.

The number of grapefruit used is: \[ \frac{7 \text{ baskets}}{1 \text{ grapefruit}} \times \frac{1 \text{ grapefruit}}{1 \text{ basket}} = 7 \text{ grapefruit} \]

Therefore the number of grapefruit remaining is 8 - 7 = 1.

The number of bananas used is: \[ \frac{7 \text{ baskets}}{1 \text{ banana}} \times \frac{3 \text{ bananas}}{1 \text{ basket}} = 21 \text{ bananas} \]

Therefore, the number of bananas remaining is 27 - 21 = 6.
This same logic can be applied to chemical reactions. Consider, for example, the combustion of acetylene (C₂H₂). The balanced equation for this reaction is shown below.

\[ 2\text{C}_2\text{H}_2(\text{l}) + 5\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \]

Suppose 75 mol C₂H₂ and 160 mol O₂ are mixed together and allowed to react. What is the maximum number of moles of CO₂ that could be produced? The first step is to decide which substance is the limiting reactant. As was done for the fruit baskets, calculate the amount of product that can form from each specified amount of reactant. The limiting reactant produces the smaller quantity, which will be the theoretical yield. All of the calculations needed are familiar ones using parts of the general solution map shown in Figure 5.4. To get the moles of CO₂ from the starting mole amounts of C₂H₂ and O₂, use moles ratios from the balanced equation.

\[
\begin{align*}
75 \text{ mol C}_2\text{H}_2 \times \frac{4 \text{ mol CO}_2}{2 \text{ mol C}_2\text{H}_2} &= 150 \text{ mol CO}_2 \\
160 \text{ mol O}_2 \times \frac{4 \text{ mol CO}_2}{5 \text{ mol O}_2} &= 130 \text{ mol CO}_2
\end{align*}
\]

Since O₂ produces the smaller amount of CO₂ it will be consumed first even though a larger number of moles of O₂ are present initially. The 130 moles of CO₂ represents the maximum yield from these reaction conditions and is the theoretical yield. You can also determine how much of the reactant in excess (C₂H₂) is left over. Calculate how many moles of C₂H₂ are needed to produce the theoretical yield and subtract this from the starting quantity.

\[
130 \text{ mol CO}_2 \times \frac{2 \text{ mol C}_2\text{H}_2}{4 \text{ mol CO}_2} = 65 \text{ mol C}_2\text{H}_2
\]

Therefore, the number of moles of C₂H₂ remaining is 75 mol - 65 mol = 10. mol. Other examples of such calculations are shown below.
Example 5.17

Problem

What is the theoretical yield (in grams) of silver if 8.75 g copper is mixed with 0.350 mol AgNO₃(aq)? The balanced equation for this reaction is shown below.

\[ \text{Cu(s)} + 2\text{AgNO₃(aq)} \rightarrow \text{Cu(NO₃)₂(aq)} + 2\text{Ag(l)} \]

Solution

What we know: g Cu; mol AgNO₃; balanced equation relating Cu, AgNO₃ and Ag

Desired answer: maximum g Ag produced

First determine the limiting reactant by calculating how many grams of silver can form from each starting amount of reactant. The solution maps for these calculations are:

\[
\begin{align*}
g \text{ Cu} & \rightarrow \text{ mol Cu} \rightarrow \text{ mol Ag} \rightarrow g \text{ Ag} \\
\text{ mol AgNO₃} & \rightarrow \text{ mol Ag} \rightarrow g \text{ Ag}
\end{align*}
\]

For the first calculation the conversion factors needed are the molar mass of Cu, the Ag/Cu mole ratio and finally the molar mass of Ag.

Putting these together yields:

\[
\begin{align*}
8.75 \text{ g Cu} & \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{2 \text{ mol Ag}}{1 \text{ mol Cu}} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 29.7 \text{ g Ag}
\end{align*}
\]

For the second calculation the conversion factors needed are the Ag/AgNO₃ mole ratio and the molar mass of Ag.

Putting these together yields:

\[
\begin{align*}
0.350 \text{ mol AgNO₃} & \times \frac{2 \text{ mol Ag}}{2 \text{ mol AgNO₃}} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 37.8 \text{ g Ag}
\end{align*}
\]

Since the starting amount of copper produces the smaller amount of silver, copper is the limiting reactant and the theoretical yield is 29.7 g Ag.
Example 5.18

Problem

How many grams of AgNO₃ remain in Example 5.17?

Solution

What we know: mol AgNO₃, g Ag produced; balanced equation relating Ag and AgNO₃

Desired answer: g AgNO₃ remaining

Since copper was found to be the limiting reactant, the AgNO₃ is in excess. Use the theoretical yield of Ag to calculate how much AgNO₃ was used and subtract this from the starting quantity. Note that the starting moles of AgNO₃ will have to be converted to grams. The solution maps for these calculations are:

\[
g \text{ Ag} \rightarrow \text{ mol Ag} \rightarrow \text{ mol AgNO₃} \rightarrow g \text{ AgNO₃ used}
\]

\[
\text{ mol AgNO₃ initially} \rightarrow g \text{ AgNO₃ initially}
\]

\[
g \text{ AgNO₃ remaining} = g \text{ AgNO₃ initially} - g \text{ AgNO₃ used}
\]

For the first calculation the conversion factors needed are the molar mass of Ag, the AgNO₃/Ag mole ratio and finally the molar mass of AgNO₃.

Putting these together yields:

\[
29.7 \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.9 \text{ g Ag}} \times \frac{2 \text{ mol AgNO₃}}{2 \text{ mol Ag}} \times \frac{169.9 \text{ g AgNO₃}}{1 \text{ mol AgNO₃}} = 46.8 \text{ g AgNO₃ used}
\]

For the second calculation the conversion factor needed is the molar mass of AgNO₃.

\[
0.350 \text{ mol AgNO₃} \times \frac{169.9 \text{ g AgNO₃}}{1 \text{ mol AgNO₃}} = 59.5 \text{ g AgNO₃ initially}
\]

Finally, the mass of AgNO₃ that remains equals 59.5 g - 46.8 g = 12.7 g.
In summary, in order to determine the theoretical yield from a given set of reaction conditions, first determine which reactant is the limiting reactant by calculating the yield from each starting amount of reactant. The reactant producing the least amount of a product is limiting and the amount of product formed from the limiting reactants is the theoretical yield.

**Check for Understanding 5.10**

1. What is the maximum number of moles of Al₂(SO₄)₃ that can be produced from the reaction of 95 moles Al(OH)₃ and 120 moles H₂SO₄? First write a balanced chemical equation for the reaction.

2. What is the theoretical yield in grams of silicon for the reaction between 2.5 kg SiCl₄ and 1.6 kg Mg?

   \[
   \text{SiCl}_4(s) + 2\text{Mg}(s) \rightarrow \text{Si}(s) + 2\text{MgCl}_2(s)
   \]

3. When a mixture of 58 g CO and 9.2 g H₂ reacts, which reactant is in excess? How many grams of the reactant in excess remain?

   \[
   \text{CO}(g) + 2\text{H}_2(g) \rightarrow \text{CH}_3\text{OH}(l)
   \]
Chapter 5 Keywords

mass spectrometry  mole  mole ratio
atomic weight  Avogadro constant  theoretical yield
atomic mass unit  molar mass  actual yield
Dalton  percent  limiting reactant
isotopes  mass percent
weighted average  stoichiometry

Supplementary Chapter 5 Check for Understanding questions

Chapter 5 Exercises

(You may use a periodic table as needed.)

1. If a single molecule of a substance weighs $7.64 \times 10^{-23}$ g, what is its molar mass?

2. Which of the following contains the smallest number of molecules?
   
   A. 1.0 g H$_2$O  
   B. 1.0 g NH$_3$  
   C. 1.0 g H$_2$  
   D. 1.0 g O$_2$  
   E. all of the above are the same

3. Ethanol (C$_2$H$_5$OH) has a density of 0.789 g/mL. How many ethanol molecules are in a teaspoon (~ 5 mL) of ethanol?

4. $10^{23}$ atoms of zinc weigh nearly the same as:
   
   A. a mole of zinc (Zn).  
   B. a half of a mole of carbon (C).  
   C. two moles of xenon (Xe).  
   D. a mole of boron (B).
5. In your own words describe what the atomic weight of an element means.

6. What is the molar mass of $C_3H_5(NO_3)_3$?

7. How many moles of ammonia are in 61 g of NH$_3$?

8. Create a solution map for calculating how many hydrogen atoms are present in a 75-g sample of CH$_3$CH$_2$OH and then make this calculation.

9. What mass of N$_2$O contains 0.50 mol of nitrogen atoms?

10. A 1979 penny is 95.05% copper by mass. If the penny weighs 3.098 g, how many copper atoms does the penny contain?

11. What is the mass percent of oxygen in K$_2$SO$_4$?

12. A substance was analyzed and found to contain only sulfur and fluorine. If there was 0.1570 g fluorine present in a 0.2011-g sample of this substance, what is the mass percent composition of this material?

13. How many moles of HNO$_3$ does a 60.0-g sample of a nitric acid solution that is 70.4% HNO$_3$ (by mass) contain?

14. How many grams of Fe$_2$O$_3$ contain 15.0 g of iron?
15. The balanced equation \(2\text{Cu}(s) + \text{O}_2(g) \rightarrow 2\text{CuO}(s)\) tells us that 1 mol of Cu

A. reacts with 1 mol \(\text{O}_2\).
B. produces 1 mol \(\text{CuO}\).
C. reacts with 32 g of \(\text{O}_2\).
D. produces 2 mol \(\text{CuO}\).
E. does both A and C.

16. Create a solution map for calculating the mass (in grams) of ammonia that can be produced from 5 kg of hydrogen gas and an excess of nitrogen according to the reaction \(\text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)\).

17. According to the reaction \(4\text{Li}(s) + \text{O}_2(g) \rightarrow 2\text{Li}_2\text{O}(s)\), how many grams of \(\text{Li}_2\text{O}\) can be produced from each gram of lithium that reacts?

18. The unbalanced equation for the reaction of \(\text{Na}_2\text{O}_2\) with water is:

\[
\text{Na}_2\text{O}_2 + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{O}_2
\]

What mass of oxygen gas will be produced if 3.25 g of sodium peroxide reacts with a large excess of water?

19. What mass of \(\text{O}_2\) is needed to react completely with 4.20 g \(\text{CH}_4\) in a combustion reaction?

20. How many grams of oxygen gas can be formed from the decomposition of 0.437 mol \(\text{HgO}\)?

21. For the reaction \(2\text{CH}_4(g) + 3\text{O}_2(g) + 2\text{NH}_3(g) \rightarrow 2\text{HCN}(g) + 6\text{H}_2\text{O}(g)\), if you start with 5.0 g of each reactant, which substance is the limiting reactant?
22. How many moles of NH₃ can be produced from the reaction of 4.00 mol N₂ and 10.0 mol H₂?

23. For the reaction \( \text{CaCO}_3(s) + 2\text{HCl(aq)} \rightarrow \text{CaCl}_2(aq) + \text{CO}_2(g) + \text{H}_2\text{O(l)} \), how many grams of CO₂ can be produced if 68.1 g CaCO₃ (molar mass = 100.09 g/mol) is mixed with 51.6 g HCl (molar mass = 36.46 g/mol)?

24. How does the actual yield of a chemical reaction compare to the theoretical yield of the reaction?

25. The unbalanced equation below describes the formation of C₂H₂.

\[
\text{CaC}_2(s) + \text{H}_2\text{O(l)} \rightarrow \text{Ca(OH)}_2(s) + \text{C}_2\text{H}_2(g)
\]

What is the maximum number of grams of C₂H₂ that can be produced if 8.00 g of CaC₂ and 8.00 g of H₂O are allowed to react?

26. When 2.34 grams of aluminum reacts with 1.50 grams of hydrochloric acid, which reactant is in excess? How much of this reactant is left over?
Chapter 5

Check for Understanding 5.1

1. The atomic weight of sodium is 22.99. What is the mass, in grams, of a billion sodium atoms?

Answer: \(3.817 \times 10^{-14} \text{ g}\)

Solution

What we know: number of Na atoms; atomic weight of Na

Desired answer: g Na

The number of sodium atoms can be converted into mass using the atomic weight of sodium. Recall that the numerical value of the atomic weight of sodium refers to the mass of a single sodium atom in atomic mass units. First, convert this value to an equivalent mass in grams.

\[
\frac{22.99 \text{ u Na}}{1 \text{ Na atom}} \times \frac{1.6605 \times 10^{-24} \text{ g Na}}{1 \text{ u Na}} = \frac{3.817 \times 10^{-23} \text{ g Na}}{\text{ Na atom}}
\]

Now use this result as the conversion factor between atoms of sodium and mass of sodium.

\[
1 \times 10^9 \text{ atoms Na} \times \frac{3.817 \times 10^{-21} \text{ g Na}}{\text{ atom Na}} = 3.817 \times 10^{-14} \text{ g Na}
\]

Check for Understanding 5.2

1. Write the solution map for the conversion of grams of sodium to atoms of sodium. Indicate the numerical ratio that is the conversion factor in each step.

Answer: The solution map is: g Na \(\rightarrow\) mol Na \(\rightarrow\) atoms Na

The conversion factor for the first step is \(\frac{1 \text{ mol Na}}{22.99 \text{ g Na}}\).

The conversion factor for the second step is \(\frac{6.022 \times 10^{23} \text{ atoms Na}}{1 \text{ mol Na}}\).
Solution

Each step of the solution map requires a conversion factor.

For \( g \text{ Na} \rightarrow \text{ mol Na} \), we need the molar mass of sodium expressed in the form \( \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \) which will cancel \( g \text{ Na} \) and introduce \( \text{ mol Na} \) into the numerator. The numerical value for the molar mass is obtained from the periodic table.

For \( \text{ mol Na} \rightarrow \text{ atoms Na} \) the conversion factor needed is the Avogadro constant expressed in the form \( \frac{6.022 \times 10^{23} \text{ atoms Na}}{1 \text{ mol Na}} \) so that \( \text{ mol Na} \) will cancel.

Check for Understanding 5.3

1. What is the mass, in grams, of one cobalt atom?

   Answer: \( 9.786 \times 10^{-23} \text{ g} \)

Solution

What we know: number of Co atoms

Desired answer: \( \text{ g Co/atom} \)

The solution map for this calculation is:

\[ \text{atom Co} \rightarrow \text{ mol Co} \rightarrow \text{ g Co} \]

The conversion factor needed in the first step is the Avogadro constant expressed in the form \( \frac{1 \text{ mol Co}}{6.022 \times 10^{23} \text{ atoms Co}} \).

The conversion factor needed in the second step is the molar mass of cobalt. The numerical value for the molar mass is obtained from the atomic weight of cobalt (58.93) and is expressed in the form \( \frac{58.93 \text{ g Co}}{1 \text{ mol Co}} \).

Putting these together yields:

\[
\frac{1 \text{ atom Co}}{6.022 \times 10^{23} \text{ atoms Co}} \times \frac{1 \text{ mol Co}}{6.022 \times 10^{23} \text{ atoms Co}} \times \frac{58.93 \text{ g Co}}{1 \text{ mol Co}} = 9.786 \times 10^{-23} \text{ g Co}
\]
2. How many atoms of aluminum are present in a piece of aluminum foil that weighs 6.27 g?

Answer: \(1.40 \times 10^{23} \text{ atoms Al}\)

**Solution**

What we know: \(g \text{ Al}\)

Desired answer: number of \(\text{Al}\) atoms

The solution map for this calculation is:

\[
\text{g Al} \rightarrow \text{mol Al} \rightarrow \text{atoms Al}
\]

The conversion factor needed in the first step is the molar mass of aluminum. The numerical value for the molar mass is obtained from the periodic table and is expressed in the form \(\frac{1 \text{ mol Al}}{26.98 \text{ g Al}}\).

The conversion factor needed in the second step is the Avogadro constant in the form \(\frac{6.022 \times 10^{23} \text{ atoms Al}}{1 \text{ mol Al}}\).

Putting these together yields:

\[
6.27 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{6.022 \times 10^{23} \text{ atoms Al}}{1 \text{ mol Al}} = 1.40 \times 10^{23} \text{ atoms Al}
\]

**Check for Understanding 5.4**

1. Calculate the molar mass of sodium phosphate.

Answer: \(163.94 \text{ g/mol}\)

**Solution**

What we know: sodium phosphate; atomic weights of elements

Desired answer: \(g \text{ sodium phosphate/mol}\)
The chemical formula for sodium phosphate is $\text{Na}_3\text{PO}_4$. The formula indicates that in 1 mole of sodium phosphate there are 3 moles of sodium, 1 mole of phosphorus and 4 moles of oxygen. From the periodic table we see that 1 mole of sodium atoms weighs 22.99 g, 1 mole of phosphorus atoms weighs 30.97 g and 1 mole of oxygen atoms weighs 16.00 g. Thus, one mole of $\text{Na}_3\text{PO}_4$ will weigh:

$$3 \text{ Na} \times \frac{22.99 \text{ g}}{1 \text{ mol Na}} = 68.97 \text{ g}$$

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

$$4 \text{ O} \times \frac{16.00 \text{ g}}{1 \text{ mol O}} = \frac{64.00 \text{ g}}{163.94 \text{ g}}$$

The molar mass of sodium phosphate is 163.94 g/mol.

**Check for Understanding 5.5**

1. Calculate the number of moles in a sample of sodium bicarbonate that weighs 6.22 g.

   **Answer:** 0.0740 mol

   **Solution**

   What we know: g sodium bicarbonate; atomic weights of elements

   Desired answer: mol sodium bicarbonate

   The solution map for this calculation is:

   $\text{g sodium bicarbonate} \rightarrow \text{mol sodium bicarbonate}$
The conversion factor needed is the molar mass of sodium bicarbonate. The chemical formula for sodium bicarbonate is NaHCO₃. From the periodic table we see that 1 mole of sodium atoms weighs 22.99 g, 1 mole of hydrogen atoms weighs 1.008 g, 1 mole of carbon atoms weighs 12.01 g and 1 mole of oxygen atoms weighs 16.00 g. Thus, one mole of NaHCO₃ will weigh:

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

$$1 \text{ H} \times \frac{1.008 \text{ g}}{1 \text{ mol H}} = 1.008 \text{ g}$$

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

$$3 \text{ O} \times \frac{16.00 \text{ g}}{1 \text{ mol O}} = 48.00 \text{ g}$$

$$\text{The molar mass of sodium bicarbonate is } 84.01 \text{ g/mol.}$$

Using this result yields:

$$\frac{6.22 \times 10²² \text{ g NaHCO}_₃}{84.01 \times 1 \text{ mol NaHCO}_₃} = 0.0740 \text{ mol NaHCO}_₃$$

2. How many molecules are present in 1.05 g of iodine (I₂)?

Answer: \[ 2.49 \times 10^{21} \text{ molecules} \]

Solution

What we know: g I₂

Desired answer: number of I₂ molecules

The solution map for this calculation is

\[ \text{g I₂} \rightarrow \text{mol I₂} \rightarrow \text{molecules I₂} \]
The conversion factor needed in the first step is the molar mass of I₂. From the periodic table we see that 1 mole of iodine atoms weighs 126.9 g. Thus, one mole of I₂ will weigh:

$$2 \text{ mol I} \times \frac{126.9 \text{ g I}}{1 \text{ mol I}} = 253.8 \text{ g}$$

This is needed in the form \(\frac{1 \text{ mol I}_2}{253.8 \text{ g I}_2}\).

The conversion factor needed in the second step is the Avogadro constant in the form \(\frac{6.022 \times 10^{23} \text{ molecules I}_2}{1 \text{ mol I}_2}\).

Putting these together yields:

$$1.05 \text{ g I}_2 \times \frac{1 \text{ mol I}_2}{253.8 \text{ g I}_2} \times \frac{6.022 \times 10^{23} \text{ molecules I}_2}{1 \text{ mol I}_2} = 2.49 \times 10^{21} \text{ molecules I}_2$$

3. What is the mass in grams of 25 molecules of CCl₂F₂?

Answer: \(5.0 \times 10^{-21} \text{ g}\)

Solution

What we know: molecules CCl₂F₂

Desired answer: g CCl₂F₂

The solution map for this calculation is:

$$\text{molecules CCl}_2\text{F}_2 \rightarrow \text{mol CCl}_2\text{F}_2 \rightarrow \text{g CCl}_2\text{F}_2$$

The conversion factor needed in the first step is the Avogadro constant in the form \(\frac{1 \text{ mol CCl}_2\text{F}_2}{6.022 \times 10^{23} \text{ molecules CCl}_2\text{F}_2}\).

The conversion factor needed in the second step is the molar mass of CCl₂F₂. From the periodic table we can get the molar masses of carbon, fluorine and chlorine and add them as follows. 12.01 g C + 2(35.45 g) Cl + 2(19.00 g) F = 120.91 g. Thus the molar mass of CCl₂F₂ is 120.91 g/mol. This is used in the form \(\frac{120.91 \text{ g CCl}_2\text{F}_2}{1 \text{ mol CCl}_2\text{F}_2}\) to convert units properly.
Putting these together yields:

\[
225 \text{ molecules } \text{CCl}_2\text{F}_2 \times \frac{1 \text{ mol } \text{CCl}_2\text{F}_2}{6.022 \times 10^{23} \text{ molecules } \text{CCl}_2\text{F}_2} \times \frac{120.91 \text{ g } \text{CCl}_2\text{F}_2}{1 \text{ mol } \text{CCl}_2\text{F}_2} = 5.0 \times 10^{21} \text{ g } \text{CCl}_2\text{F}_2
\]

**Check for Understanding 5.6**

1. What is the numerical ratio between the number of hydrogen atoms and oxygen atoms in \((\text{NH}_4)_3\text{PO}_4\)?

   **Answer:** \(\frac{12 \text{ H atoms}}{4 \text{ O atoms}}\)

   **Solution**

   In each formula unit of \((\text{NH}_4)_3\text{PO}_4\) there are 12 H atoms and 4 O atoms.

2. What is the numerical ratio between the number of moles of sulfur and moles of \(\text{Al}_2(\text{SO}_4)_3\)?

   **Answer:** \(\frac{3 \text{ mol S}}{1 \text{ mol } \text{Al}_2(\text{SO}_4)_3}\)

   **Solution**

   There are 3 S atoms in each formula unit of \(\text{Al}_2(\text{SO}_4)_3\). This ratio can be converted to a ratio of moles by multiplying each quantity by the Avogadro constant.

**Check for Understanding 5.7**

1. Calculate the number of moles of barium nitrate that contain 0.81 mol of oxygen atoms.

   **Answer:** 0.14 mol

   **Solution**

   What we know: mol oxygen atoms

   Desired answer: mol barium nitrate
The solution map for this calculation is:

\[
\text{mol oxygen atoms} \rightarrow \text{mol barium nitrate}
\]

The chemical formula for barium nitrate is \( \text{Ba(NO}_3\text{)}_2 \). The formula indicates that in 1 mole of barium nitrate there are 6 moles of oxygen atoms. This relationship can be applied as follows.

\[
0.81 \text{ mol O} \times \frac{1 \text{ mol Ba(NO}_3\text{)}_2}{6 \text{ mol O}} = 0.14 \text{ mol Ba(NO}_3\text{)}_2
\]

2. How many carbon atoms are present in 16.0 g \( \text{C}_3\text{H}_8 \)?

Answer: \( 6.56 \times 10^{23} \) atoms

Solution

What we know: \( \text{g C}_3\text{H}_8 \)

Desired answer: number of carbon atoms

The solution map for this calculation is:

\[
\text{g C}_3\text{H}_8 \rightarrow \text{mol C}_3\text{H}_8 \rightarrow \text{mol C} \rightarrow \text{atoms C}
\]

The conversion factor needed in the first step is the molar mass of \( \text{C}_3\text{H}_8 \). From the periodic table we can get the molar masses of carbon and hydrogen and add them as follows. \( 3(12.01 \text{ g}) \text{ C} + 8(1.008 \text{ g}) \text{ H} = 44.09 \text{ g} \). Thus the molar mass of \( \text{C}_3\text{H}_8 \) is 40.09 g/mol. This is needed in the form \( \frac{1 \text{ mol C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} \).

The formula of \( \text{C}_3\text{H}_8 \) indicates that in 1 mole of \( \text{C}_3\text{H}_8 \) there are 3 moles of carbon atoms. This relationship can be applied as the conversion factor in the second step in the form \( \frac{3 \text{ mol C}}{1 \text{ mol C}_3\text{H}_8} \).

The conversion factor needed in the last step is the Avogadro constant in the form \( \frac{6.022 \times 10^{23} \text{ atoms C}}{1 \text{ mol C}} \).

Putting these together yields:

\[
16.0 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} \times \frac{3 \text{ mol C}}{1 \text{ mol C}_3\text{H}_8} \times \frac{6.022 \times 10^{23} \text{ atoms C}}{1 \text{ mol C}} = 6.56 \times 10^{23} \text{ atoms C}
\]
3. How many grams of silver are present in 25.0 g AgBr?

Answer: 14.4 g

Solution

What we know: g AgBr

Desired answer: g Ag

The solution map for this calculation is:

\[ \text{g AgBr} \rightarrow \text{mol AgBr} \rightarrow \text{mol Ag} \rightarrow \text{g Ag} \]

The conversion factor needed in the first step is the molar mass of AgBr. From the periodic table we can get the molar masses of silver and bromine and add them as follows. 107.9 g Ag + 79.90 g Br = 187.8 g. Thus the molar mass of AgBr is 187.8 g/mol. This is needed in the form \( \frac{1\text{mol AgBr}}{187.8\text{g AgBr}} \).

The formula of AgBr indicates that in 1 mole of AgBr there is 1 mole of silver. This relationship can be applied as the conversion factor in the second step in the form \( \frac{1\text{mol Ag}}{1\text{mol AgBr}} \).

The conversion factor needed in the last step is the molar mass of Ag expressed in the form \( \frac{107.9\text{g Ag}}{1\text{mol Ag}} \).

Putting these together yields:

\[
25.0 \text{ g AgBr} \times \frac{1\text{mol AgBr}}{187.8 \text{g AgBr}} \times \frac{1\text{mol Ag}}{1\text{mol AgBr}} \times \frac{107.9\text{g Ag}}{1\text{mol Ag}} = 14.4 \text{ g Ag}
\]

Check for Understanding 5.8

1. What is the mass percent of carbon in C₆H₁₂O₆?

Answer: 40.00 %

Solution

What we know: C₆H₁₂O₆; atomic weights of elements

Desired answer: mass % for C in C₆H₁₂O₆
Since the mass percent values apply to any sample of C₆H₁₂O₆, it is convenient to consider one mole of this compound. From the periodic table we can get the molar masses of carbon, hydrogen and oxygen and add them as follows. 6(12.01 g) C + 12(1.008 g) H + 6(16.00 g) O = 180.16 g. Thus the molar mass of C₆H₁₂O₆ is 180.16 g/mol, so 180.16 g represents the total mass of material. The formula indicates that 1 mole of C₆H₁₂O₆ contains 6 moles of carbon. Therefore the mass of carbon (component mass) present is:

\[
6 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 72.06 \text{ g C}
\]

These values are used to obtain the mass percent for carbon.

\[
\text{mass\% C} = \frac{72.06 \text{ g C}}{180.16 \text{ g C}_6\text{H}_12\text{O}_6} \times 100 = 40.00\%
\]

**Check for Understanding 5.9**

1. Based on the reaction between iron(III) oxide and carbon, how many moles of carbon are needed to produce 1.0 kmol of iron?

\[
\text{Fe}_2\text{O}_3(s) + 3\text{C(s)} \rightarrow 2\text{Fe(s)} + 3\text{CO(g)}
\]

Answer: 1500 mol

**Solution**

What we know: kmol Fe; balanced equation relating Fe and C

Desired answer: mol C

The solution map for this problem is:

\[
\text{kmol Fe} \rightarrow \text{mol Fe} \rightarrow \text{mol C}
\]

The conversion factor needed in the first step is that between kmol and mol in the form \( \frac{10^3 \text{ mol Fe}}{1 \text{ kmol Fe}} \).

The conversion factor needed in the second step is the mole ratio for these two substances from the balanced equation in the form \( \frac{3 \text{ mol C}}{2 \text{ mol Fe}} \).
Putting these together yields:

\[
1.0 \text{ kmol Fe} \times \frac{10^3 \text{ mol Fe}}{1 \text{ kmol Fe}} \times \frac{3 \text{ mol C}}{2 \text{ mol Fe}} = 1500 \text{ mol C}
\]

2. What mass of NaHCO₃ will react with 0.00257 mol HCl?

\[
\text{NaHCO}_3(s) + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{CO}_3(aq)
\]

Answer: 0.216 g

Solution

What we know: mol HCl; balanced equation relating HCl and NaHCO₃

Desired answer: g NaHCO₃

The solution map for this problem is:

\[
\text{mol HCl} \rightarrow \text{mol NaHCO}_3 \rightarrow \text{g NaHCO}_3
\]

The conversion factor needed in the first step is the mole ratio for these two substances from the balanced equation in the form \( \frac{1 \text{ mol NaHCO}_3}{1 \text{ mol HCl}} \).

The conversion factor needed in the second step is the molar mass of NaHCO₃ expressed in the form \( \frac{84.01 \text{ g NaHCO}_3}{1 \text{ mol NaHCO}_3} \).

Putting these together yields:

\[
0.00257 \text{ mol HCl} \times \frac{1 \text{ mol NaHCO}_3}{1 \text{ mol HCl}} \times \frac{84.01 \text{ g NaHCO}_3}{1 \text{ mol NaHCO}_3} = 0.216 \text{ g NaHCO}_3
\]

3. How many grams of sodium metal are need to produce 125 g hydrogen gas by the reaction below?

\[
2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + \text{H}_2(\text{g})
\]

Answer: \(2.85 \times 10^3\) g
Solution

What we know: \( g \text{ H}_2 \); balanced equation relating \( \text{H}_2 \) and \( \text{Na} \)

Desired answer: \( g \text{ Na} \)

The solution map for this problem is:

\[ g \text{ H}_2 \rightarrow \text{mol H}_2 \rightarrow \text{mol Na} \rightarrow g \text{ Na} \]

The conversion factor needed in the first step is the molar mass of \( \text{H}_2 \) expressed in the form \( \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \).

The conversion factor needed in the second step is the mole ratio for these two substances from the balanced equation in the form \( \frac{2 \text{ mol Na}}{1 \text{ mol H}_2} \).

The conversion factor needed in the last step is the molar mass of \( \text{Na} \) expressed in the form \( \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} \).

Putting these together yields:

\[
125 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{2 \text{ mol Na}}{1 \text{ mol H}_2} \times \frac{22.99 \text{ g Na}}{1 \text{ mol Na}} = 2.85 \times 10^3 \text{ g Na}
\]

Check for Understanding 5.10

1. What is the maximum number of moles of \( \text{Al}_2(\text{SO}_4)_3 \) that can be produced from the reaction of 95 moles \( \text{Al(OH)}_3 \) and 120 moles \( \text{H}_2\text{SO}_4 \)? First write a balanced chemical equation for the reaction?

Answer: 40, mol

Solution

The balanced equation for the reaction is:

\[
2\text{Al(OH)}_3(s) + 3\text{H}_2\text{SO}_4(aq) \rightarrow \text{Al}_2(\text{SO}_4)_3(aq) + 6\text{H}_2\text{O}(l)
\]
What we know: $\text{mol Al(OH)₃}$; $\text{mol H₂SO₄}$; balanced equation relating $\text{Al(OH)₃}$, $\text{H₂SO₄}$ and $\text{Al₂(SO₄)₃}$

Desired answer: maximum moles $\text{Al₂(SO₄)₃}$ produced

First determine the limiting reactant by calculating how many moles of $\text{Al₂(SO₄)₃}$ can form from each starting amount of reactant. The solution maps for these calculations are:

$$\text{mol Al(OH)₃} \rightarrow \text{mol Al₂(SO₄)₃}$$

$$\text{mol H₂SO₄} \rightarrow \text{mol Al₂(SO₄)₃}$$

For the first calculation the conversion factor needed is the $\text{Al₂(SO₄)₃}/\text{Al(OH)₃}$ mole ratio from the balanced equation.

Applying this yields:

$$95 \text{ mol Al(OH)₃} \times \frac{1 \text{ mol Al₂(SO₄)₃}}{2 \text{ mol Al(OH)₃}} = 48 \text{ mol Al₂(SO₄)₃}$$

For the second calculation the conversion factor needed is the $\text{Al₂(SO₄)₃}/\text{H₂SO₄}$ mole ratio from the balanced equation.

Applying this yields:

$$120 \text{ mol H₂SO₄} \times \frac{1 \text{ mol Al₂(SO₄)₃}}{3 \text{ mol H₂SO₄}} = 40 \text{ mol Al₂(SO₄)₃}$$

Since the starting amount of $\text{H₂SO₄}$ produces the smaller amount of $\text{Al₂(SO₄)₃}$, sulfuric acid is the limiting reactant and the maximum moles of $\text{Al₂(SO₄)₃}$ is 40. moles.

2. What is the theoretical yield in grams of silicon for the reaction between 2.5 kg $\text{SiCl₄}$ and 1.6 kg $\text{Mg}$?

$$\text{SiCl₄(s)} + 2\text{Mg(s)} \rightarrow \text{Si(s)} + 2\text{MgCl₂(s)}$$

Answer: $4.1 \times 10^2$ g

Solution

What we know: $\text{kg SiCl₄}$; $\text{kg Mg}$; balanced equation relating $\text{SiCl₄}$, $\text{Mg}$ and $\text{Si}$

Desired answer: maximum g $\text{Si}$ produced
First determine the limiting reactant by calculating how many grams of silicon can form from each starting amount of reactant. The solution maps for these calculations are:

\[
\text{kg SiCl}_4 \rightarrow \text{g SiCl}_4 \rightarrow \text{mol SiCl}_4 \rightarrow \text{mol Si} \rightarrow \text{g Si}
\]

\[
\text{kg Mg} \rightarrow \text{g Mg} \rightarrow \text{mol Mg} \rightarrow \text{mol Si} \rightarrow \text{g Si}
\]

For the first calculation the conversion factors needed are that between kg and g, the molar mass of SiCl\(_4\), the Si/SiCl\(_4\) mole ratio and finally the molar mass of Si.

Putting these together yields:

\[
2.5 \text{ kg SiCl}_4 \times \frac{10^3 \text{ g SiCl}_4}{1 \text{ kg SiCl}_4} \times \frac{1 \text{ mol SiCl}_4}{169.89 \text{ g SiCl}_4} \times \frac{1 \text{ mol Si}}{1 \text{ mol SiCl}_4} \times \frac{28.09 \text{ g Si}}{1 \text{ mol Si}} = 4.1 \times 10^2 \text{ g Si}
\]

For the second calculation the conversion factors needed are that between kg and g, the molar mass of Mg, the Si/Mg mole ratio and finally the molar mass of Si.

Putting these together yields:

\[
2.5 \text{ kg Mg} \times \frac{10^3 \text{ g Mg}}{1 \text{ kg Mg}} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{1 \text{ mol Si}}{2 \text{ mol Mg}} \times \frac{28.09 \text{ g Si}}{1 \text{ mol Si}} = 9.2 \times 10^2 \text{ g Si}
\]

Since the starting amount of SiCl\(_4\) produces the smaller amount of silicon, SiCl\(_4\) is the limiting reactant and the theoretical yield is 4.1 \times 10^2 \text{ g Si}.

3. When a mixture of 58 g CO and 9.2 g H\(_2\) reacts, which reactant is in excess? How many grams of the reactant in excess remain?

\[
\text{CO(g)} + 2\text{H}_2(\text{g}) \rightarrow \text{CH}_3\text{OH(1)}
\]

Answers:

H\(_2\) 0.7 g

Solution

What we know: g CO; g H\(_2\); balanced equation relating CO, H\(_2\) and CH\(_3\)OH

Desired answer: g excess reactant remaining
To find which reactant is in excess, determine the limiting reactant by calculating how many moles of CH$_3$OH can form from each starting amount of reactant. The solution maps for these calculations are:

\[
g \text{CO} \rightarrow \text{mol CO} \rightarrow \text{mol CH}_3\text{OH}
\]
\[
g \text{H}_2 \rightarrow \text{mol H}_2 \rightarrow \text{mol CH}_3\text{OH}
\]

For the first calculation the conversion factors needed are the molar mass of CO and the CH$_3$OH/CO mole ratio.

Putting these together yields:

\[
58 \text{ g CO} \times \frac{1 \text{ mol CO}}{28.01 \text{ g CO}} \times \frac{1 \text{ mol CH}_3\text{OH}}{1 \text{ mol CO}} = 2.1 \text{ mol CH}_3\text{OH}
\]

For the second calculation the conversion factors needed are the molar mass of H$_2$ and the CH$_3$OH/H$_2$ mole ratio.

Putting these together yields:

\[
9.2 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{1 \text{ mol CH}_3\text{OH}}{2 \text{ mol H}_2} = 2.3 \text{ mol CH}_3\text{OH}
\]

Since the starting amount of CO produces the smaller amount of CH$_3$OH, CO is the limiting reactant and H$_2$ is the reactant in excess.

Now use the theoretical yield of CH$_3$OH to calculate how much H$_2$ was used and subtract this from the starting quantity. The solution maps for these calculations are:

\[
\text{mol CH}_3\text{OH} \rightarrow \text{mol H}_2 \rightarrow \text{g H}_2 \text{ used}
\]
\[
g \text{H}_2 \text{ remaining} = \text{g H}_2 \text{ initially} - \text{g H}_2 \text{ used}
\]

For the first calculation the conversion factors needed are the H$_2$/CH$_3$OH mole ratio and the molar mass of H$_2$.

Putting these together yields:

\[
2.1 \text{ mol CH}_3\text{OH} \times \frac{2 \text{ mol H}_2}{1 \text{ mol CH}_3\text{OH}} \times \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} = 8.5 \text{ g H}_2 \text{ used}
\]

Therefore, the mass of H$_2$ that remains equals 9.2 g - 8.5 g = 0.7 g.
Chapter 5

1. 46.0 g/mol

2. B

3. $5 \times 10^{22}$ molecules

4. D

6. 227.10 g/mol

7. 3.6 mol

8. \[ \text{g CH}_3\text{CH}_2\text{OH} \rightarrow \text{mol CH}_3\text{CH}_2\text{OH} \rightarrow \text{mol H} \rightarrow \text{atoms H} \]
   \[5.9 \times 10^{24} \text{ atoms}\]

9. 11 g

10. $2.790 \times 10^{22} \text{ atoms}$

11. 36.72%

12. 21.93% S
    78.07% F

13. 0.670 mol

14. 21.4 g

15. B
16. \( \text{kg H}_2 \rightarrow \text{g H}_2 \rightarrow \text{mol H}_2 \rightarrow \text{mol NH}_3 \rightarrow \text{g NH}_3 \)

17. 2.152 g

18. 0.667 g

19. 16.8 g

20. 6.83 g

21. \( \text{O}_2 \)

22. 6.67 mol

23. 29.9 g

24. The actual yield is always less than the theoretical yield.

25. 5.78 g

26. \( \text{Al is in excess} \)

1.97 g remain