## Concepts and Skills Review

This review section summarizes many of the concepts that you encountered in your Grade 11 chemistry course. You may wish to read through this section before continuing with the rest of the textbook, to remind yourself of important terms, equations, and calculations. Sample problems and practice problems are included to help you review your skills. As well, you can refer to the information in this section, as needed, while you work through the textbook. Table R. 1 lists the topics that are covered in the Concepts and Skills Review.

Table R. 1 Topics Included in Concepts and Skills Review

| R.1 | Matter | R. 9 | Types of Chemical Reactions |
| :--- | :--- | :--- | :--- |
| R.2 | Representing Atoms and Ions | R.10 | Ionic Equations |
| R.3 | The Periodic Table | R.11 | Mole Calculations |
| R.4 | Chemical Bonds | R.12 | Concentration Calculations |
| R.5 | Representing Molecules | R.13 | Stoichiometric Calculations |
| R.6 | Naming Binary Compounds | R.14 | Representing Organic Molecules |
| R. 7 | Writing Chemical Formulas | R.15 | Isomers of Organic Compounds |
| R.8 | Balancing Chemical Formulas |  |  |

## R. 1 Matter

Chemistry is the study of the properties and changes of matter. Matter is defined as anything that has mass and takes up space. All matter can be classified into two groups: pure substances and mixtures.

A pure substance has a definite chemical composition. Examples of pure substances are carbon dioxide, $\mathrm{CO}_{2}$, and nitrogen, $\mathrm{N}_{2}$. Pure substances can be further classified into elements and compounds.

- An element is a pure substance that cannot be separated chemically into any simpler substances. Oxygen gas, $\mathrm{O}_{2(\mathrm{~g})}$, solid carbon, $\mathrm{C}_{(\mathrm{s})}$, and copper metal, $\mathrm{Cu}_{(\mathrm{s})}$, are examples of elements.
- A compound is a pure substance that results when two or more elements combine chemically to form a different substance. Water, $\mathrm{H}_{2} \mathrm{O}_{(\ell)}$, and salt, $\mathrm{NaCl}_{(\mathrm{s})}$, are two examples of compounds.

A mixture is a physical combination of two or more kinds of matter. Each component in a mixture retains its identity. There are two kinds of mixtures: heterogeneous mixtures and homogeneous mixtures.

- In a heterogeneous mixture, the different components are clearly visible. A mixture of sand and table salt is a heterogeneous mixture, with small grains of both sand and salt visible.
- In a homogeneous mixture (also called a solution) the components are blended together so that the mixture looks like a single substance. A spoonful of table salt in a glass of water results in a homogeneous mixture, since the salt dissolves in the water to produce a salt-water solution.


## R. 2 Representing Atoms and Ions

An atom is the smallest particle of an element that still retains the identity and properties of the element. An atom is composed of one or more protons, neutrons, and electrons. Each atom of an element has the same number of protons in its nucleus.

- The atomic number (symbol $\mathbf{Z}$ ) of an element is the number of protons in the nucleus of each atom of the element.
- The mass number (symbol $\mathbf{A}$ ) is the total number of protons and neutrons in the nucleus of each atom. If two atoms of an element have the same number of protons, but different numbers of neutrons, they are called isotopes. Isotopes have the same atomic number but different mass numbers.
- The atomic symbol is different for each element. Some examples of atomic symbols are O for oxygen, Au for gold, and Br for bromine.
Figure R. 1 summarizes the notation that is used to express the atomic number, mass number, and atomic symbol of an element.

When an atom gains or loses electrons, it becomes an ion. Atoms and ions can be represented by Bohr-Rutherford diagrams or by Lewis structures.

A Bohr-Rutherford diagram shows the number of electrons in each energy level. Figure R. 2 shows Bohr-Rutherford diagrams for atoms and ions of magnesium and fluorine.


Figure R. 2 Bohr-Rutherford diagrams of (A) a magnesium atom, (B) a magnesium ion, (C) a fluorine atom, and (D) a fluoride ion

In a Lewis structure of an element, the dots show the number of electrons in the outer, valence shell. The symbol represents the nucleus and the inner energy levels. Figure R. 3 shows Lewis structures for atoms and ions of magnesium and fluoride.
A
$\dot{\mathrm{Mg}}$.
B
$\mathrm{Mg}^{2+}$


Figure R. 3 Lewis structures of (A) a magnesium atom, (B) a magnesium ion, (C) a fluorine atom, and (D) a fluoride ion


## Z

atomic number
Figure R. 1 The mass number is written at the top left of the atomic symbol. The atomic number is written at the bottom left.

## R. 3 The Periodic Table

The elements can be organized according to similarities in their properties and atomic structure. The periodic table is a system for organizing the elements, by atomic number, into groups (columns) and periods (rows). In the periodic table, elements with similar properties are in the same column. The names that are associated with some sections of the periodic table are shown in Figure R.4.


- Elements are arranged in seven numbered periods (horizontal rows) and 18 numbered groups (vertical columns).
- Groups are numbered according to two different systems. The current system numbers the group from 1 to 18. An older system numbers the groups from I to VIII, and separates them into two categories labelled A and B .
- The elements in the eight A groups are the main-group elements. They are also called
the representative elements. The elements in the ten B groups are known as the transition elements.
- Group 1 (IA) elements are known as alkali metals. They react with water to form alkaline, or basic, solutions.
- Group 2 (IIA) elements are known as alkaline earth metals. They react with oxygen to form compounds called oxides, which react with water to form alkaline solutions. Early chemists called all metal oxides "earths."
- Group 17 (VIIA) elements are known as halogens, from the Greek word hals, meaning "salt." Elements in this group combine with other elements to form compounds called salts.
- Group 18 (VIIIA) elements are known as noble gases. Noble gases do not combine naturally with any other elements.

The periodic law states that when the elements are arranged in order of increasing atomic number, a regular repetition of properties is observed. This statement can be used to predict trends in the properties of the elements. Figure R. 5 summarizes the periodic trends of four properties of atoms: atomic size, ionization energy, electron affinity, and electronegativity. These four properties affect the structure of molecules and ions, and they are key to understanding the properties of matter.


A Trends in atomic size


Trends in ionization energy


C Trends in electron affinity


## R. 4 Chemical Bonds

Bonding involves interactions between the valence electrons of atoms.
Valence electrons are electrons that occupy the outer energy level of an atom. Bonding usually follows the octet rule, which states that atoms attain a more stable electron configuration with eight electrons in their valence shell.

- When electron(s) are transferred between a metal and a non-metal, the electrostatic attraction between the positive metal ion (cation) and the negative non-metal ion (anion) is called an ionic bond. Figure R. 6 shows the formation of an ionic bond between the metal calcium and the non-metal fluorine.


[^0]

Figure R. 7 Lewis structure showing the covalent bonding in a molecule of $\mathrm{CH}_{4}$
(A) $\quad \bullet: 10$
( $\quad: \mathrm{O}=0^{\circ}$.
Figure R. 10 Lewis structures of $\mathbf{0}_{2}$

- When a pair of electrons is shared between two non-metal atoms, the attraction is called a covalent bond. A single covalent bond involves one pair of electrons shared between two atoms. A double bond involves two pairs of electrons shared between two atoms, and a triple bond involves three pairs of electrons. Figure R. 7 shows the formation of a covalent bond between two non-metals, carbon and hydrogen atoms, in a molecule of methane, $\mathrm{CH}_{4}$.

The electronegativity of an element is a relative measure of the ability of its atoms to attract electrons in a chemical bond. The periodic table in Appendix C gives the electronegativities of the elements. The difference in the electronegativities ( $\triangle E N$ ) of two atoms is used to predict the type of bond that will form. By convention, $\triangle E N$ is always positive. Therefore, always subtract the smaller electronegativity from the larger one. Figure R. 8 illustrates the range in bond character for different values of $\Delta E N$.


Figure R. 8 Chemical bonds range from mostly ionic to mostly covalent.
A polar covalent bond is a covalent bond between atoms that have different electronegativities. The electron pair is unevenly shared in a polar covalent bond. Therefore, the atom with the higher electronegativity has a slight negative charge and the atom with the lower electronegativity has a slight positive charge. This separation of charges, or polarity, is shown using the Greek letter delta, $\delta$. Figure R. 9 gives examples of an ionic bond, a polar covalent bond, and a mostly covalent bond.

| A | B | C |
| :---: | :---: | :---: |
| NaCl | $\begin{aligned} & \delta+\quad \delta- \\ & \mathrm{H}-\mathrm{Cl} \end{aligned}$ | $\mathrm{Br}-\mathrm{Cl}$ |
| $\Delta E N=3.16-0.93=2.23$ | $\Delta E N=3.16-2.20=0.96$ | $\Delta E N=3.16-2.96=0.20$ |

Figure R.9 (A) ionic bond in NaCl , (B) polar covalent bond in HCl , and (C) covalent bond in BrCl

## R. 5 Representing Molecules

The Lewis structure of a molecule shows the electrons that are in the valence shells of all the atoms in the molecule. A common variation of a Lewis structure uses a dash (the symbol of a single bond) to represent one shared pair of electrons. Figure R.10(A) shows a Lewis structure for a molecule of $\mathrm{O}_{2}$ that contains a double covalent bond. Figure R.10(B) shows a similar structure for the same molecule, in which the electrons of the double bond have been replaced by the symbol of a double bond.

## R. 6 Naming Binary Compounds

Two non-metals can combine to form a binary compound: a compound with only two kinds of atoms. The less electronegative element is usually written on the left, and the more electronegative element is usually written on the right. For example, sulfur and oxygen can combine to form $\mathrm{SO}_{2}$ and $\mathrm{SO}_{3}$. Carbon and chlorine can combine to form $\mathrm{CCl}_{4}$.

## Naming a Covalent Binary Compound

To name a covalent binary compound, follow these steps.
Step 1 Give the first atom its full name. For example, use sulfur for $\mathrm{SO}_{2}$ and carbon for $\mathrm{CCl}_{4}$.
Step 2 Give the second atom its ion name, with the suffix -ide. For example, use oxide for $\mathrm{SO}_{2}$, and chloride for $\mathrm{CCl}_{4}$.
Step 3 Use a prefix to indicate the number of each type of atom. Table R. 2 lists the prefixes that are commonly used to designate the number of each type of atom in a molecular compound.
Step 4 Put the parts of the name together. Thus, $\mathrm{SO}_{2}$ is named sulfur dioxide, and $\mathrm{CCl}_{4}$ is named carbon tetrachloride.

A metal and a non-metal can combine to form an ionic binary compound. In an ionic binary compound, the cation is written on the left and the anion is written on the right. For example, sodium and chlorine can combine to form NaCl . Calcium and chloride can combine to form $\mathrm{CaCl}_{2}$.

## Naming an Ionic Binary Compound

To name an ionic binary compound, follow the steps below.
Step 1 Give the metal ion its full name. For example, use sodium for NaCl and calcium for $\mathrm{CaCl}_{2}$.
Step 2 Give the non-metal ion its ion name, with the suffix -ide. For example, use chloride for NaCl and $\mathrm{CaCl}_{2}$.
Step 3 Put the parts of the name together. Thus, NaCl is named sodium chloride, and $\mathrm{CaCl}_{2}$ is named calcium chloride. Note that prefixes are not used for ionic binary compounds. That is, calcium chloride is not named calcium dichloride, even though two chloride ions are present for each calcium ion.

## The Classical System and the Stock System

Transition metals may have more than one valence. Two systems are available to name ionic compounds that contain transition metals.

- For the classical system, use the suffix -ous to indicate the metal ion with the lower valence. Use the suffix -ic to indicate the metal ion with the higher valence. The earliest discovered elements are sometimes named using Latin names. For example, FeO is named ferrous oxide, and $\mathrm{Fe}_{2} \mathrm{O}_{3}$ is named ferric oxide. Other common examples are stannous oxide, SnO , and mercuric nitride, $\mathrm{Hg}_{3} \mathrm{~N}_{2}$.
- The Stock system was devised by the German chemist Alfred Stock. For the Stock system, use Roman numerals to indicate the valence of the metal cation. Place a Roman numeral in brackets after the name of the first element. Examples are copper(II) oxide, CuO, and copper(I) oxide, $\mathrm{Cu}_{2} \mathrm{O}$.


## R. 7 Writing Chemical Formulas

A chemical formula indicates the type and number of atoms that are present in a compound. You can write the chemical formula of a compound by using the valence of each type of atom. When you write the chemical formula of a neutral compound that contains ions, the sum of the positive valences plus the negative valences must equal zero. This statement is known as the zero sum rule. To write a chemical formula, follow the steps below.

## Writing a Chemical Formula

Step 1 Write the unbalanced formula, placing the element or ion with a positive valence first.
Step 2 Write the valence of each element on top of the appropriate symbol. The names and formulas of commonly used ions are listed in Appendix E.
Step 3 Cross over the numerical value of each valence and write this number as the subscript for the other element or ion in the compound. Do not include negative or positive signs, and do not include the subscript 1 in the formula. Bracket the formula for a polyatomic ion if its subscript is greater than 1.
Step 4 If necessary, divide through by any common factor. (Use your knowledge of chemical bonding to decide when to divide by a common factor. For example, the formula for hydrogen peroxide, $\mathrm{H}_{2} \mathrm{O}_{2}$, should not be reduced further, since the formula HO cannot represent a neutral molecule.)

The following Sample Problem shows you how to use these steps to write a chemical formula. Note that the steps are listed for only the first two solutions.

## Sample Problem

Writing a Chemical Formula From the Name of a Compound

## Problem

Write the chemical formula of each compound.
(a) magnesium fluoride
(b) zinc telluride
(c) aluminum carbonate
(d) ammonium phosphite
(e) $\operatorname{tin}(I V)$ sulfate
(f) ferrous oxalate

## What Is Required?

Use the name of each compound to write a chemical formula for the compound.

## What Is Given?

You are given the names of the compounds. From each name, you can identify the types of atoms that are present in the compound.

## Plan Your Strategy

Use the periodic table to find the valence of each atom in the name. (Valence numbers usually correspond to the common ion charge. If the compound includes a transition metal with more than one valence, the name of the compound will indicate which valence is used.) Then follow the steps you have just learned to write each chemical formula.

## Act on Your Strategy

(a) The valences are +2 for Mg and -1 for $F$.

Step 1 Magnesium has a positive valence, so write the unbalanced formula as MgF .

Step 2 Write the valences above each element.

$$
\begin{aligned}
& +2-1 \\
& \mathrm{Mg} \mathrm{~F}
\end{aligned}
$$

Step 3 Cross over the numerical value of each valence.

$$
\mathrm{MgF}_{2}
$$

Step 4 Since there is only one Mg atom, you do not need to divide by a common factor.
(c) $+3-2$

Al CO 3
$\mathrm{Al}_{2}\left(\mathrm{CO}_{3}\right)_{3}$
(e) $+4-2$

Sn SO 4
$\mathrm{Sn}_{2}\left(\mathrm{SO}_{4}\right)_{4}=\mathrm{Sn}\left(\mathrm{SO}_{4}\right)_{2}$
(b) The valences are +2 for Zn , and -2 for Te .

Step 1 Zinc has a positive valence, so write the unbalanced formula as ZnTe .

Step 2 Write the valences above each element.

$$
\begin{aligned}
& +2-2 \\
& \mathrm{Zn} \mathrm{Te}
\end{aligned}
$$

Step 3 Cross over the numerical value of each valence.

$$
\mathrm{Zn}_{2} \mathrm{Te}_{2}
$$

Step 4 Divide by the common factor, 2, to give the chemical formula ZnTe .
(d) $+1-3$
$\mathrm{NH}_{4} \mathrm{PO}_{3}$
$\left(\mathrm{NH}_{4}\right)_{3} \mathrm{PO}_{3}$
(f) $+2-2$
$\mathrm{Fe} \mathrm{C}_{2} \mathrm{O}_{4}$
$\mathrm{Fe}_{2}\left(\mathrm{C}_{2} \mathrm{O}_{4}\right)_{2}=\mathrm{FeC}_{2} \mathrm{O}_{4}$

## Check Your Solution

If you have time, use each chemical formula to name the compound. Then check that your name and the original name match.

## Practice Problems

1. Name each compound.
(a) $\mathrm{XeF}_{4}$
(d) $\mathrm{SF}_{6}$
(g) $\mathrm{CaCO}_{3}$
(b) $\mathrm{PCl}_{5}$
(e) $\mathrm{N}_{2} \mathrm{O}_{4}$
(h) $\mathrm{BaS}_{2} \mathrm{O}_{3}$
(c) CO
(f) NaF
(i) $\mathrm{NI}_{3}$
2. Give both the classical name and the Stock name of each compound.
(a) $\mathrm{TiO}_{2}$
(c) $\mathrm{NiBr}_{2}$
(b) $\mathrm{CoCl}_{2}$
(d) HgO
3. Write the chemical formula of each compound.
(a) copper(II) hydroxide
(g) iron(III) acetate
(b) mercuric nitride
(h) potassium peroxide
(c) manganese(II) hydrogen carbonate
(i) ammonium dichromate
(d) sulfur hexabromide
(j) stannous permanganate
(e) potassium chromate
(k) plumbous nitrate
(f) arsenic trichloride
4. Write the correct chemical formula of each compound.
(a) zinc perchlorate
(h) lead(II) bicarbonate
(b) mercury(II) nitride
(i) silver oxalate
(c) stannous fluoride
(j) platinum(IV) oxide
(d) ammonium dihydrogen phosphite
(k) silicon carbide
(e) manganese(IV) silicate
(I) nickel(III) sulfite
(f) ferric hydroxide
(m) tin(II) carbonate
(g) cobalt(III) nitride
( $n$ ) aluminum permanganate

## R. 8 Balancing Chemical Equations

A chemical equation shows the reactants (the starting materials) and the products (the new materials that form) in a chemical change. Chemical equations are balanced to reflect the fact that atoms are conserved in a chemical reaction. The basic process of balancing an equation involves trial and error-going back and forth between reactants and products to find the correct balance. The systematic approach that is outlined below can be helpful.

## Balancing a Chemical Equation

Step 1 List all the atomic species that are involved in the equation to determine which element(s) are not balanced.
Step 2 First balance the most complex substance: the compound that contains the most kinds or the largest number of atoms. Balance the atoms that occur in the largest numbers. Leave hydrogen, oxygen, and elements that occur in smaller numbers until later.
Step 3 Balance any polyatomic ions that occur on both sides of the equation as one unit, rather than as separate atoms.
Step 4 Balance any hydrogen or oxygen atoms that occur in a combined and uncombined state.
Step 5 Balance any other element that occurs in its uncombined state.

## Sample Problem

## Balancing a Chemical Equation

Problem
Balance the following chemical equation.
$\mathrm{NH}_{3(\mathrm{~g})}+\mathrm{H}_{2} \mathrm{O}(\ell)+\mathrm{Y}_{2}\left(\mathrm{SO}_{4}\right)_{3(\mathrm{aq})} \rightarrow\left(\mathrm{NH}_{4}\right)_{2} \mathrm{SO}_{4(\mathrm{qq})}+\mathrm{Y}(\mathrm{OH})_{3(\mathrm{~s})}$

## Solution

Step 1 List the atoms on each side of the equation.

| 1 | $\mid$ | N | $\mid$ | 2 |
| ---: | :---: | :---: | :---: | ---: |
| 2 | $\mathrm{\mid}$ | Y | \| | 1 |
| 3 | $\mid$ | $\mathrm{SO}_{4}$ | $\mid$ | 1 |
| 5 | $\mid$ | H | $\mid$ | 11 |
| 13 | $\mid$ | O | \| | 7 |

Steps 2 and $3 \mathrm{SO}_{4}{ }^{2-}$ is a polyatomic ion, and it is present in the most complex substance. It appears on both sides of the equation, so it can be balanced as a unit. Balance $\mathrm{SO}_{4}{ }^{2-}$ by putting a 3 on the right side, in front of $\left(\mathrm{NH}_{4}\right)_{2} \mathrm{SO}_{4 \text { (aq) }}$.
$\mathrm{NH}_{3(\mathrm{~g})}+\mathrm{H}_{2} \mathrm{O}_{(\ell)}+\mathrm{Y}_{2}\left(\mathrm{SO}_{4}\right)_{3(\text { aq })} \rightarrow 3\left(\mathrm{NH}_{4}\right)_{2} \mathrm{SO}_{4(\mathrm{aq})}+\mathrm{Y}(\mathrm{OH})_{3(\mathrm{~s})}$
Since there are now 6 N atoms on the right side, put a 6 in front of $\mathrm{NH}_{3(\mathrm{~g})}$ on the left side. Balance Y by putting a 2 in front of $\mathrm{Y}(\mathrm{OH})_{3(\mathrm{~s})}$ on the right side.
$6 \mathrm{NH}_{3(\mathrm{~g})}+\mathrm{H}_{2} \mathrm{O}_{(\ell)}+\mathrm{Y}_{2}\left(\mathrm{SO}_{4}\right)_{3(\text { aq })} \rightarrow 3\left(\mathrm{NH}_{4}\right)_{2} \mathrm{SO}_{4(\mathrm{aq})}+2 \mathrm{Y}(\mathrm{OH})_{3(\mathrm{~s})}$
Step 4 There are now 6 O atoms on the right side and 1 O atom on the left side. (This does not include O atoms in the $\mathrm{SO}_{4}{ }^{2-}$ units, which are already balanced.) Put a 6 in front of $\mathrm{H}_{2} \mathrm{O}$ on the left side. This also balances H .

$$
6 \mathrm{NH}_{3(\mathrm{~g})}+6 \mathrm{H}_{2} \mathrm{O}_{(\ell)}+\mathrm{Y}_{2}\left(\mathrm{SO}_{4}\right)_{3(\mathrm{aq})} \rightarrow 3\left(\mathrm{NH}_{4}\right)_{2} \mathrm{SO}_{4(\mathrm{aq})}+2 \mathrm{Y}(\mathrm{OH})_{3(\mathrm{~s})}
$$

## Practice Problems

5. Balance each equation.
(a) $\mathrm{H}_{2} \mathrm{O}_{(\ell)}+\mathrm{NO}_{2(\mathrm{~g})} \rightarrow \mathrm{HNO}_{3(\text { aq })}+\mathrm{NO}_{(\mathrm{g})}$
(b) $\mathrm{NaOH}_{(\text {aq) }}+\mathrm{Cl}_{2(\mathrm{~g})} \rightarrow \mathrm{NaCl}_{(\mathrm{aq})}+\mathrm{NaClO}_{(\mathrm{aq})}+\mathrm{H}_{2} \mathrm{O}_{(\ell)}$
(c) $\left(\mathrm{NH}_{4}\right)_{3} \mathrm{PO}_{4(\text { aq })}+\mathrm{Ba}(\mathrm{OH})_{2(\text { aq })} \rightarrow \mathrm{Ba}_{3}\left(\mathrm{PO}_{4}\right)_{2(\mathrm{~s})}+\mathrm{NH}_{4} \mathrm{OH}_{(\text {aq })}$
(d) $\mathrm{Li}_{(\mathrm{s})}+\mathrm{H}_{2} \mathrm{O}_{(\ell)} \rightarrow \mathrm{LiOH}_{(\mathrm{aq})}+\mathrm{H}_{2(\mathrm{~g})}$
(e) $\mathrm{NH}_{3(\mathrm{~g})}+\mathrm{O}_{2(\mathrm{~g})} \rightarrow \mathrm{N}_{2(\mathrm{~g})}+\mathrm{H}_{2} \mathrm{O}_{(\ell)}$
(f) $\mathrm{Sb}_{4} \mathrm{~S}_{6(\mathrm{~s})}+\mathrm{O}_{2(\mathrm{~g})} \rightarrow \mathrm{Sb}_{4} \mathrm{O}_{6(\mathrm{~s})}+\mathrm{SO}_{2(\mathrm{~g})}$
(g) $\mathrm{Al}\left(\mathrm{NO}_{3}\right)_{3 \text { (aq) }}+\mathrm{H}_{2} \mathrm{SO}_{4 \text { (aq) }} \rightarrow \mathrm{Al}_{2}\left(\mathrm{SO}_{4}\right)_{3(\text { aq })}+\mathrm{HNO}_{3 \text { (aq) }}$
(h) $\mathrm{Cu}\left(\mathrm{NO}_{3}\right)_{2(\mathrm{~s})} \rightarrow \mathrm{CuO}_{(\mathrm{s})}+\mathrm{NO}_{2(\mathrm{~g})}+\mathrm{O}_{2(\mathrm{~g})}$
(i) $\mathrm{KI}_{(\mathrm{aq})}+\mathrm{MnO}_{2(\mathrm{~s})}+\mathrm{H}_{2} \mathrm{SO}_{4(\text { aq })} \rightarrow \mathrm{K}_{2} \mathrm{SO}_{4(\text { aq })}+\mathrm{MnSO}_{4(\mathrm{aq})}+\mathrm{H}_{2} \mathrm{O}_{(\ell)}+\mathrm{I}_{2(\mathrm{~s})}$
(j) $\mathrm{C}_{6} \mathrm{H}_{6(\ell)}+\mathrm{O}_{2(\mathrm{~g})} \rightarrow \mathrm{CO}_{2(\mathrm{~g})}+\mathrm{H}_{2} \mathrm{O}_{(\mathrm{g})}$

## R. 9 Types of Chemical Reactions

Most chemical reactions can be classified as one of four main types of reactions. As you can see in Table R.3, these four types of reactions are classified by counting the number of reactants and products.

- In a synthesis reaction, two or more reactants combine to produce a single, different substance.
- In a decomposition reaction, a compound breaks down into elements or simpler compounds.
- In a single displacement reaction, one element in a compound is replaced by another element.
- In a double displacement reaction, the cations of two ionic compounds exchange places, resulting in the formation of two new compounds.

Table R. 3 Four Types of Reactions

| Name | General form | Example |
| :--- | :--- | :--- |
| synthesis reaction | $\mathrm{A}+\mathrm{B} \rightarrow \mathrm{AB}$ | $2 \mathrm{Cu}_{(\mathrm{s})}+\mathrm{S}_{(\mathrm{g})} \rightarrow \mathrm{Cu}_{2} \mathrm{~S}_{(\mathrm{s})}$ |
| decomposition reaction | $\mathrm{AB} \rightarrow \mathrm{A}+\mathrm{B}$ | $2 \mathrm{HgO}_{(\mathrm{g})} \rightarrow 2 \mathrm{Hg}_{(\ell)}+\mathrm{O}_{2(\mathrm{~g})}$ |
| single displacement reaction | $\mathrm{A}+\mathrm{BC} \rightarrow \mathrm{B}+\mathrm{AC}$ | $\mathrm{Zn}_{(\mathrm{s})}+\mathrm{CuCl}_{2(\mathrm{aq})} \rightarrow \mathrm{Cu}_{(\mathrm{s})}+\mathrm{ZnCl}_{2(\mathrm{aq})}$ |
| double displacement reaction | $\mathrm{AB}+\mathrm{CD} \rightarrow \mathrm{CB}+\mathrm{AD}$ | $2 \mathrm{NaOH}_{(\mathrm{aq})}+\mathrm{CuSO}_{4(\mathrm{aq})} \rightarrow \mathrm{Na}_{2} \mathrm{SO}_{4(\mathrm{aq})}+\mathrm{Cu}(\mathrm{OH})_{2(\mathrm{~s})}$ |

Many combustion reactions do not fit into any of the four categories in Table R.3. Therefore, combustion reactions are usually classified separately. A combustion reaction occurs when a compound reacts in the presence of oxygen to form oxides, heat, and light (burning). For example, sulfur reacts with oxygen to produce sulfur dioxide.

$$
\mathrm{S}_{8(\mathrm{~s})}+8 \mathrm{O}_{2(\mathrm{~g})} \rightarrow 8 \mathrm{SO}_{2(\mathrm{~g})}
$$

Combustion reactions are common for compounds that are composed of carbon and hydrogen atoms. These compounds are called hydrocarbons.

- Complete combustion of a hydrocarbon occurs when the hydrocarbon reacts completely in the presence of sufficient oxygen. The complete combustion of a hydrocarbon produces only water vapour and carbon dioxide gas as products. The complete combustion of propane, $\mathrm{C}_{3} \mathrm{H}_{8}$, is shown below.

$$
\mathrm{C}_{3} \mathrm{H}_{8(\mathrm{~g})}+5 \mathrm{O}_{2(\mathrm{~g})} \rightarrow 3 \mathrm{CO}_{2(\mathrm{~g})}+4 \mathrm{H}_{2} \mathrm{O}_{(\mathrm{g})}
$$

- Incomplete combustion of a hydrocarbon occurs when there is not enough oxygen present for the hydrocarbon to react completely. The incomplete combustion of a hydrocarbon produces water, carbon dioxide, carbon monoxide, and solid carbon in varying amounts. More than one balanced equation is possible for the incomplete combustion of a hydrocarbon. One possible equation for the incomplete combustion of propane is shown below.

$$
2 \mathrm{C}_{3} \mathrm{H}_{8(\mathrm{~g})}+7 \mathrm{O}_{2(\mathrm{~g})} \rightarrow 2 \mathrm{CO}_{2(\mathrm{~g})}+8 \mathrm{H}_{2} \mathrm{O}_{(\mathrm{g})}+2 \mathrm{CO}_{(\mathrm{g})}+2 \mathrm{C}_{(\mathrm{s})}
$$

## R. 10 Ionic Equations

When an ionic compound dissolves, the ions break away from their crystal lattice and become mobile in the solution. This process is called dissociation. You can summarize the dissociation process using a dissociation equation, as illustrated by the following examples.

$$
\begin{aligned}
& \mathrm{BaSO}_{4(\mathrm{~s})} \rightarrow \mathrm{Ba}^{2+}{ }_{\text {(aq) }}+\mathrm{SO}_{4}^{2-}{ }_{\text {(aq) }} \\
& \mathrm{Al}\left(\mathrm{NO}_{3}\right)_{3(\mathrm{~s})} \rightarrow \mathrm{Al}^{3+}{ }_{\text {(aq) }}+3 \mathrm{NO}_{3}^{--}{ }_{\text {(aq) }}
\end{aligned}
$$

Notice that, in addition to being balanced by atoms, these dissociation equations are also balanced by charge. The total net charge is zero on both sides of each equation.

For reactions that occur between ionic compounds in aqueous solution, a total ionic equation is used to show all the ions that are present and any ions that combine to form a precipitate. The total ionic equation for the reaction of silver nitrate and potassium chromate is given below.
$2 \mathrm{Ag}^{+}{ }_{\text {(aq) }}+2 \mathrm{NO}_{3}^{-}{ }_{\text {(aq) }}+2 \mathrm{~K}^{+}{ }_{\text {(aq) }}+\mathrm{CrO}_{4}{ }^{2-}{ }_{\text {(aq) }} \rightarrow 2 \mathrm{~K}^{+}{ }_{\text {(aq) }}+2 \mathrm{NO}_{3}{ }^{-}{ }_{\text {(aq) }}+\mathrm{Ag}_{2} \mathrm{CrO}_{4(\mathrm{~s})}$
A net ionic equation shows only the ions that react and the precipitate that forms. The ions that do not participate in the reaction are called spectator ions, and they are omitted from the net ionic equation. In the reaction of silver nitrate and potassium chromate, $\mathrm{K}^{+}{ }_{\text {(aq) }}$ and $\mathrm{NO}_{3}{ }^{-}{ }_{(\text {aq })}$ are the spectator ions. They are not included in the net ionic equation for the reaction.

$$
2 \mathrm{Ag}^{+}{ }_{(\mathrm{aq})}+\mathrm{CrO}_{4}{ }^{2-}{ }_{(\text {aq })} \rightarrow \mathrm{Ag}_{2} \mathrm{CrO}_{4(\mathrm{~s})}
$$

## Practice Problems

6. What type of reaction is represented by each balanced equation?
(a) $2 \mathrm{~N}_{2} \mathrm{O}_{(\mathrm{g})} \rightarrow 2 \mathrm{~N}_{2(\mathrm{~g})}+\mathrm{O}_{2(\mathrm{~g})}$
(b) $\mathrm{NO}_{2(\mathrm{~g})}+\mathrm{NO}_{2(\mathrm{~g})} \rightarrow \mathrm{N}_{2} \mathrm{O}_{4(\mathrm{~g})}$
(c) $8 \mathrm{Al}_{(\mathrm{s})}+3 \mathrm{Co}_{3} \mathrm{O}_{4(\mathrm{~s})} \rightarrow 9 \mathrm{Co}_{(\mathrm{s})}+4 \mathrm{Al}_{2} \mathrm{O}_{3(\mathrm{~s})}$
7. Write the dissociation equation for each compound.
(a) $\mathrm{MgSO}_{3}$
(c) $\mathrm{Sn}\left(\mathrm{SO}_{4}\right)_{2}$
(e) $\mathrm{Al}_{2}\left(\mathrm{C}_{2} \mathrm{O}_{4}\right)_{3}$
(b) $\mathrm{Fe}(\mathrm{OH})_{2}$
(d) $\mathrm{Ag}_{2} \mathrm{~S}$
8. Aqueous solutions of several pairs of ionic solids are prepared and then mixed together. For each pair listed below, write

- the dissociation equation that occurs when each solid dissolves in water
- the balanced, total ionic equation
- the net ionic reaction, if a precipitate forms
- the spectator ions
(a) lead nitrate + zinc iodide
(b) potassium chlorate + calcium chloride
(c) ammonium phosphate + copper(II) chlorate
(d) sodium hydroxide + iron(III) nitrate
(e) calcium nitrate + potassium carbonate


## R. 11 Mole Calculations

The mole (symbol mol) is the SI base unit for quantity of matter. One mole contains the same number of particles as exactly 12 g of carbon- 12 . When measured experimentally, one mole is the quantity of matter that contains $6.022 \times 10^{23}$ items. This number is called the Avogadro constant ( $\boldsymbol{N}_{\mathrm{A}}$ ). The following concept organizer summarizes the possible conversions between the number of moles ( $n$ ), mass ( $m$ ), molar mass ( $M$ ), number of particles ( $N$ ), and the Avogadro constant ( $N_{\mathrm{A}}$ ).

## Concept Organizer

## Conversions Used in Mole Calculations

Unit analysis $\quad \mathrm{mol}=\mathrm{g} \times \frac{\mathrm{mol}}{\mathrm{g}} \quad \mathrm{mol}=$ number of particles $\times \frac{\mathrm{mol}}{6.02 \times 10^{23} \text { particles }}$


Unit analysis $\quad \mathrm{g}=\mathrm{mol} \times \frac{\mathrm{g}}{\mathrm{mol}} \quad$ number of particles $=\mathrm{mol} \times \frac{6.02 \times 10^{23} \text { particles }}{\mathrm{mol}}$

## Sample Problem

## Calculating Mass from Number of Particles

## Problem

What is the mass of $2.734 \times 10^{24}$ formula units of $\mathrm{Cu}_{3}\left(\mathrm{PO}_{4}\right)_{2}$ ?

## What Is Required?

You need to calculate the mass of $2.734 \times 10^{24}$ formula units of $\mathrm{Cu}_{3}\left(\mathrm{PO}_{4}\right)_{2}$.

## What Is Given?

From the periodic table, you can obtain the molar masses of each element in the compound.
You know that one mole contains $6.022 \times 10^{23}$ particles.

## Plan Your Strategy

Step 1 Use the molar masses of the elements to calculate the molar mass of each formula unit.

Step 2 Divide the molar mass of one formula unit by the Avogadro constant to obtain the mass of each formula unit in grams. Multiply this value by the number of formula units to obtain the total mass in grams. These calculations can be performed as a single step.

## Act on Your Strategy

Step 1 Molar mass of $\mathrm{Cu}_{3}\left(\mathrm{PO}_{4}\right)_{2}=380.59 \mathrm{~g} / \mathrm{mol}$
Step 2 Total mass ( g ) $=2.734 \times 10^{24}$ formula units

$$
\begin{aligned}
& \times \frac{1 \mathrm{~mol}}{6.022 \times 10^{23} \text { formula units }} \times \frac{380.59 \mathrm{~g}}{\mathrm{mot}} \\
& =1.728 \times 10^{3} \mathrm{~g}
\end{aligned}
$$

## Check Your Solution

Make sure that the units cancel out to give the correct unit for the answer.

## Practice Problems

9. What is the mass of $3.04 \times 10^{-4} \mathrm{~mol}$ of baking soda, $\mathrm{NaHCO}_{3}$ ?
10. How many moles of iron(III) acetate have a mass of $1.36 \times 10^{3} \mathrm{~g}$ ?
11. How many moles of chlorine atoms are present in a $9.2 \times 10^{-2} \mathrm{~g}$ sample of zirconium(IV) chloride, $\mathrm{ZrCl}_{4}$ ?
12. What mass of sodium carbonate decahydrate contains $5.47 \times 10^{23}$ atoms of oxygen?
13. A 5.00 g sample is $88.4 \%$ zinc hydroxide. How many atoms of zinc are in this sample?
14. What mass of carbon is found in $8.25 \times 10^{-6} \mathrm{~mol}^{\mathrm{m}} \mathrm{CaC}_{2}$ ?
15. Sample A is $90.2 \% \mathrm{Fe}_{3} \mathrm{O}_{4}$ and has a mass of 4.82 g . Sample B is $100.0 \%$ ferric hydroxide and has a mass of 6.0 g . Show, by calculation, which sample contains more atoms of iron.
16. A sample of coal is $2.81 \%$ sulfur by mass. How many moles of sulfur are present in 5.00 t of the sample?

## R. 12 Concentration Calculations

The concentration of a solution is an expression of the amount of solute that is present in a given volume of solution. (A solute is a substance that is dissolved in a solution.) Table R. 4 lists some units of concentration.

Table R. 4 Summary of Units of Concentration

| Unit of concentration | Mathematical expression |
| :---: | :---: |
| mass/volume percent (m/v) | $\frac{\text { mass of solute }(\mathrm{g})}{\text { volume of solution }} \times 100 \%$ |
| mass/mass percent (m/m) | $\frac{\text { mass of solute }(\mathrm{g})}{\text { mass of solution }(\mathrm{g})} \times 100 \%$ |
| volume/volume percent (v/v) | $\frac{\text { volume of solute }(\mathrm{mL})}{\text { volume of solution }(\mathrm{mL})} \times 100 \%$ |
| parts per million (ppm) | $\frac{\text { mass of solute (g) }}{\text { mass of solution (g) }} \times 10^{6} \quad \frac{\mathrm{mg}}{\mathrm{L}}$ in $\mathrm{H}_{2} \mathrm{O}$ |
| parts per billion (ppb) | $\frac{\text { mass of solute (g) }}{\text { mass of solution (g) }} \times 10^{9} \quad \frac{\mu \mathrm{~g}}{\mathrm{~L}}$ in $\mathrm{H}_{2} \mathrm{O}$ |
| molar concentration (mol/L), C | $\frac{\text { mol of solute, } n}{\text { volume of solution, } V} \text {, also written as } \mathrm{C}=\frac{n}{V}$ |

## Sample Problem

## Calculating the Concentration of a Solution

## Problem

What volume of $0.0259 \mathrm{~mol} / \mathrm{L}$ iron(III) nitrate can be prepared from 30.00 g of $\mathrm{Fe}\left(\mathrm{NO}_{3}\right)_{3} \cdot 9 \mathrm{H}_{2} \mathrm{O}$ ?

## What Is Required?

You need to calculate the volume of $0.0259 \mathrm{~mol} / \mathrm{L}$ iron(III) nitrate solution that can be made using 30.00 g of $\mathrm{Fe}\left(\mathrm{NO}_{3}\right)_{3} \cdot 9 \mathrm{H}_{2} \mathrm{O}$.

## What Is Given?

The molar concentration of the iron(III) nitrate solution is $0.0259 \mathrm{~mol} / \mathrm{L}$. The mass of $\mathrm{Fe}\left(\mathrm{NO}_{3}\right)_{3} \cdot 9 \mathrm{H}_{2} \mathrm{O}$ is 30.00 g .

## Plan Your Strategy

Step 1 Use the chemical formula $\mathrm{Fe}\left(\mathrm{NO}_{3}\right)_{3} \cdot 9 \mathrm{H}_{2} \mathrm{O}$ to calculate the molar mass.
Step 2 Use the molar mass to calculate the number of moles of $\mathrm{Fe}\left(\mathrm{NO}_{3}\right)_{3} \cdot 9 \mathrm{H}_{2} \mathrm{O}$ that are present in 30.00 g .
Step 3 Use the number of moles of $\mathrm{Fe}\left(\mathrm{NO}_{3}\right)_{3} \cdot 9 \mathrm{H}_{2} \mathrm{O}$ and the molar concentration of the solution of iron(III) nitrate to calculate the volume of the solution. Use the equation $C=\frac{n}{V}$, where $C$ is the molar concentration, $n$ is the number of moles, and $V$ is the volume.

## Act on Your Strategy

Step 1 The molar mass of $\mathrm{Fe}\left(\mathrm{NO}_{3}\right)_{3} \cdot 9 \mathrm{H}_{2} \mathrm{O}$ is $403.97 \mathrm{~g} / \mathrm{mol}$.
Step 2 Number of moles of $\mathrm{Fe}\left(\mathrm{NO}_{3}\right)_{3} \cdot 9 \mathrm{H}_{2} \mathrm{O}$

$$
\begin{aligned}
& =30.00 \mathrm{~g} \mathrm{Fe}\left(\mathrm{NO}_{3}\right)_{3} \cdot 9 \mathrm{H}_{2} \mathrm{O} \times \frac{1 \mathrm{~mol}}{403.97 \mathrm{~g}} \\
& =7.426 \times 10^{-2} \mathrm{~mol}
\end{aligned}
$$

Step 3 Volume of solution $(V)=\frac{n}{C}$

$$
\begin{aligned}
& =\frac{7.426 \times 10^{-2} \mathrm{~mol}}{0.0259 \mathrm{~mol} / \mathrm{L}} \\
& =2.87 \mathrm{~L}
\end{aligned}
$$

## Check Your Solution

The units cancel out to give an answer in litres. The smallest number of significant digits in the question is three, so the answer is also given to three significant digits.

## Practice Problems

17. A salt solution has a concentration of $1.00 \mathrm{~mol} / \mathrm{L}$. What volume of this solution is needed to prepare 2.00 L of a solution that has a concentration of $0.655 \mathrm{~mol} / \mathrm{L}$ ?
18. A 10.00 g sample of $\mathrm{CaCl}_{2}$ is added to water to make 100.0 mL of solution. Then a 400.0 mL sample of water is added to this solution. Determine the concentration of $\mathrm{Cl}^{-}$ions in the diluted solution.
19. A 50.0 mL sample of $0.85 \mathrm{~mol} / \mathrm{L} \mathrm{NaHCO}_{3}$ is diluted to a volume of 250.0 mL . Then a 50.0 mL sample of this dilute solution is evaporated to dryness. What mass of $\mathrm{NaHCO}_{3}$ remains?
20. What volume of $0.502 \mathrm{~mol} / \mathrm{L} \mathrm{KOH}$ solution must be diluted to prepare 1.500 L of $0.100 \mathrm{~mol} / \mathrm{L} \mathrm{KOH}$ ?
21. A 500.0 mL sample of a $1.02 \times 10^{-4} \mathrm{~mol} / \mathrm{L}$ lead(II) acetate solution evaporates to dryness. What mass of $\mathrm{Pb}\left(\mathrm{C}_{2} \mathrm{H}_{3} \mathrm{O}_{2}\right)_{2}$ remains?
22. A 13.6 g sample of NaCl and a 7.34 g sample of $\mathrm{CaCl}_{2}$ are dissolved in water to make 200.0 mL of solution. What is the concentration of $\mathrm{Cl}^{-}$in this solution?
23. A 50.0 g sample of $\mathrm{Al}\left(\mathrm{NO}_{3}\right)_{3}$ is dissolved in water to prepare 1500.0 mL of solution. What is the concentration, in $\mathrm{mol} / \mathrm{L}$, of $\mathrm{NO}_{3}{ }^{-}$ions in the solution?
24. What condition must exist for the concentration of a solution expressed as $\mathrm{m} / \mathrm{m}$ percent to be the same as its concentration expressed as $\mathrm{m} / \mathrm{v}$ percent?
25. A sample of lead nitrate, with a mass of 0.00372 g , is completely dissolved in 250.0 mL of water. Assume that no change in volume occurs. Calculate the following concentrations.
(a) the concentration of the solution, expressed in mol/L
(b) the concentration of $\mathrm{Pb}^{2+}$, expressed in ppm
(c) the concentration of the solution, expressed as $\mathrm{m} / \mathrm{m}$ percent

## R. 13 Stoichiometric Calculations

Stoichiometry is the study of the relative quantities of reactants and products in a chemical reaction. Calculations in stoichiometry usually involve mole ratios. A mole ratio is a ratio that compares the number of moles of different substances in a balanced chemical equation. For example, Table R. 5 shows the formation of ammonia, $\mathrm{NH}_{3}$, from nitrogen, $\mathrm{N}_{2}$, and hydrogen, $\mathrm{H}_{2}$. Suppose that you are given the number of moles of nitrogen that react with hydrogen. Using the mole ratio of nitrogen to ammonia, you can predict how many moles of ammonia will be produced.

Table R. 5 The Formation of Ammonia

| $\mathrm{N}_{2(\mathrm{~g})}$ | $3 \mathrm{H}_{2(\mathrm{~g})}$ | 2NH3(g) |
| :---: | :---: | :---: |
| 1 mol | 3 mol | 2 mol |
| If you are given 0.5 mol | Then $0.5 \mathrm{~mol} \mathrm{~N}_{2} \times \frac{3 \mathrm{~mol} \mathrm{H}_{2}}{1 \mathrm{~mol} \mathrm{~N}_{2}}=1.5 \mathrm{~mol} \mathrm{H}_{2}$ | And $0.5 \mathrm{~mol} \mathrm{~N} 2 \times \frac{3 \mathrm{~mol} \mathrm{NH}_{3}}{1 \mathrm{~mol} \mathrm{~N}_{2}}=1.0 \mathrm{~mol} \mathrm{NH}_{3}$ |

Stoichiometric problems can usually be solved by following the steps on the next page. Table R.6, on the next page, lists some equations that are useful for stoichiometric calculations.

## Solving a Stoichiometric Problem

Step 1 Write a balanced chemical equation.
Step 2 If you are given the mass, number of particles, or volume of a substance, convert this value to the number of moles.
Step 3 Calculate the number of moles of the required substance based on the number of moles of the given substance, using the appropriate mole ratio.

Step 4 Convert the number of moles of the required substance to the appropriate unit, as directed by the question.

Table R. 6 Stoichiometric Calculations

## CONCEPT CHECK

Table R. 6 includes the ideal gas law, $P V=n R T$. Practise manipulating this equation to solve for each of the four variables, and for the gas constant $R$.

| Given | Equation |
| :--- | :---: |
| mass of a reactant or product $(m)$ and molar mass $(M)$ | $m=n \times M$ or $n=\frac{m}{M}$ |
| volume of gas $(V)$ at the known temperature $(T)$ <br> and pressure $(P)$ | $P V=n R T$ or $n=\frac{P V}{R T}$ |
| volume of solution $(V)$ of known molar <br> concentration $(C)$ | $C=\frac{n}{V}$ or $n=C \times V$ |

## Sample Problem

## Mass and Particle Stoichiometry

## Problem

Passing chlorine gas through molten sulfur produces liquid disulfur dichloride. How many molecules of chlorine react to produce 50.0 g of disulfur dichloride?

## What Is Required?

You need to determine the number of molecules of chlorine gas that produce 50.0 g of disulfur dichloride.

## What Is Given?

Reactant: chlorine, $\mathrm{Cl}_{2}$
Reactant: sulfur, S
Product: disulfur dichloride, $\mathrm{S}_{2} \mathrm{Cl}_{2} \rightarrow 50.0 \mathrm{~g}$

## Plan Your Strategy

Follow the steps for solving stoichiometric problems.

## Act on Your Strategy

Step 1 Write the balanced chemical equation.

$$
\mathrm{Cl}_{2(\mathrm{~g})}+2 \mathrm{~S}_{(\ell)} \rightarrow \mathrm{S}_{2} \mathrm{Cl}_{2(\ell)}
$$

Step 2 Convert the number of grams of the product to moles.

$$
\frac{50.0 \mathrm{~g} \mathrm{~S}_{2} \mathrm{Cl}_{2}}{135 \mathrm{~g} / \mathrm{mol}}=0.370 \mathrm{~mol} \mathrm{~S} \mathrm{~S}_{2} \mathrm{Cl}_{2}
$$

Step 3 Use the mole ratio to calculate the amount of chlorine.

$$
\begin{aligned}
& \frac{\text { Amount Cl }}{2} \\
& 0.370 \mathrm{~mol} \mathrm{~S}_{2} \mathrm{Cl}_{2}=\frac{1 \mathrm{~mol} \mathrm{Cl}_{2}}{1 \mathrm{~mol} \mathrm{~S}_{2} \mathrm{Cl}_{2}} \\
&\left(0.370 \mathrm{~mol} \mathrm{~S}_{2} \mathrm{Cl}_{2}\right) \frac{\mathrm{Amount} \mathrm{Cl}_{2}}{0.370 \mathrm{~mol} \mathrm{~S}_{2} \mathrm{Cl}_{2}}=\left(0.370 \mathrm{~mol} \mathrm{~S}_{2} \mathrm{Cl}_{2}\right) \frac{1 \mathrm{~mol} \mathrm{Cl}_{2}}{1 \mathrm{~mol} \mathrm{~S}_{2} \mathrm{Cl}_{2}} \\
&{\mathrm{Amount} \mathrm{Cl}_{2}}^{=} 0.370 \mathrm{~mol} \mathrm{Cl}_{2}
\end{aligned}
$$

Step 4 Convert the number of moles of chlorine gas to the number of particles.
$0.370 \mathrm{~mol} \mathrm{Cl}_{2} \times 6.02 \times 10^{23}$ molecules $/ \mathrm{mol}$

$$
=2.23 \times 10^{23} \text { molecules } \mathrm{Cl}_{2}
$$

Therefore, $2.23 \times 10^{23}$ molecules of chlorine react to produce 50.0 g of disulfur dichloride.

## Check Your Solution

The units are correct. $2.0 \times 10^{23}$ is about one third of a mole, or 0.33 mol . One third of a mole of disulfur dichloride has a mass of 45 g , which is close to 50 g . The answer is reasonable.

## Sample Problem

## Calculating the Limiting Reagent and Percentage Yield

## Problem

A 18.9 g sample of Cu and a 82.0 mL sample of $16 \mathrm{~mol} / \mathrm{L}_{\mathrm{HNO}}^{3}$ are allowed to react. The balanced chemical equation is given below.
$\mathrm{Cu}_{(\mathrm{s})}+4 \mathrm{HNO}_{3(\text { aq })} \rightarrow \mathrm{Cu}\left(\mathrm{NO}_{3}\right)_{2(\text { aq })}+2 \mathrm{H}_{2} \mathrm{O}_{(\ell)}+2 \mathrm{NO}_{2(\mathrm{~g})}$
(a) Determine the mass of $\mathrm{NO}_{2}$ that could be produced.
(b) If 22.6 g of $\mathrm{NO}_{2}$ is actually produced in this reaction, calculate the percentage yield.

## What Is Required?

Predict the mass of $\mathrm{NO}_{2}$ that should be produced in the reaction. Next, calculate the percentage yield for this reaction, given the mass of $\mathrm{NO}_{2}$ that is actually produced.

## What Is Given?

You know the mass of each reactant: 18.9 g Cu and 82.0 mL of $16 \mathrm{~mol} / \mathrm{L}$ $\mathrm{HNO}_{3}$.

## Plan Your Strategy

(a) Determine the mass of $\mathrm{NO}_{2}$. Calculate the number of moles of each reactant that is present. Determine which reactant will be used up first, that is, which is the limiting reactant. Use the limiting reactant and the mole ratio to determine the number of moles of $\mathrm{NO}_{2}$ produced. Convert the number of moles of $\mathrm{NO}_{2}$ to grams.
(b) Calculate the percentage yield, using the following equation.

Percentage yield $=\frac{\text { Actual yield }}{\text { Theoretical yield }} \times 100 \%$

## Act on Your Strategy

(a) Determine the mass of $\mathrm{NO}_{2}$.

Number of moles of $\mathrm{Cu}=18.9 \mathrm{~g} \mathrm{Cu} \times \frac{1 \mathrm{~mol}}{63.55 \mathrm{~g}}$

$$
=0.297 \mathrm{~mol}
$$

Number of moles of $\mathrm{HNO}_{3}=C \times V$

$$
\begin{aligned}
& =16.0 \frac{\mathrm{~mol}}{\mathrm{~L}} \times 0.0820 \mathrm{~L} \\
& =1.31 \mathrm{~mol}
\end{aligned}
$$

$0.297 \mathrm{~mol} \mathrm{Cu} \times \frac{4 \mathrm{~mol} \mathrm{HNO}_{3}}{1 \mathrm{~mol} \mathrm{Cu}}=1.19 \mathrm{~mol} \mathrm{HNO}_{3}$
If all the Cu reacts, 1.19 mols of $\mathrm{HNO}_{3}$ will react. More than this amount of $\mathrm{HNO}_{3}$ is given ( 1.31 mol ). Therefore, $\mathrm{HNO}_{3}$ is in excess and Cu is the limiting reactant.
$0.297 \mathrm{~mol} \mathrm{Cu} \times \frac{2 \mathrm{~mol} \mathrm{NO}_{2}}{1 \mathrm{~mol} \mathrm{Cu}}=0.594 \mathrm{~mol} \mathrm{NO}_{2}$
$0.584 \mathrm{~mol} \mathrm{NO}_{2} \times \frac{46.01 \mathrm{~g}}{1 \mathrm{~mol}}=27.3 \mathrm{~g} \mathrm{NO}_{2}$
Theoretically, a 27.3 g sample of $\mathrm{NO}_{2}$ is produced.
(b) Calculate the percentage yield.

$$
\begin{aligned}
\text { Percentage yield } & =\frac{22.6 \mathrm{~g}}{27.3 \mathrm{~g}} \times 100 \% \\
& =82.8 \%
\end{aligned}
$$

## Check Your Solution

The units are correct. A 27.3 g sample of $\mathrm{NO}_{2}$ is predicted, but only a 22.6 g sample is produced. Therefore, it makes sense that the percentage yield is $82.8 \%$.

## Practice Problems

26. In the smelting process, iron(II) sulfide is converted to iron(III) oxide. The balanced chemical equation is given below.
$4 \mathrm{FeS}_{(\mathrm{s})}+7 \mathrm{O}_{2(\mathrm{~g})} \rightarrow 2 \mathrm{Fe}_{2} \mathrm{O}_{3(\mathrm{~s})}+4 \mathrm{SO}_{2(\mathrm{~g})}$
Calculate the mass of $\mathrm{Fe}_{2} \mathrm{O}_{3(\mathrm{~s})}$ that is produced when 37.62 g of FeS and 22.56 g of $\mathrm{O}_{2}$ are allowed to react.
27. 40.00 mL of $0.0256 \mathrm{~mol} / \mathrm{L}$ gold(III) chloride is treated with 85.00 mL of $0.105 \mathrm{~mol} / \mathrm{L}$ potassium iodide.
$\mathrm{AuCl}_{3(\mathrm{aq})}+3 \mathrm{KI}_{\text {(aq) }} \rightarrow \mathrm{AuI}_{(\mathrm{s})}+3 \mathrm{KCl}_{(\mathrm{aq})}+\mathrm{I}_{2 \text { (aq) }}$
What is the theoretical yield of $\mathrm{AuI}_{(\mathrm{s})}$ produced?
28. When 300.0 mL of $\mathrm{TiCl}_{4(\mathrm{~g})}$, at $48.0^{\circ} \mathrm{C}$ and a pressure of 105.3 kPa , is reacted with 0.4320 g of magnesium, 0.4016 g of titanium is produced.
$\mathrm{TiCl}_{4(\mathrm{~g})}+2 \mathrm{Mg}_{(\mathrm{s})} \rightarrow \mathrm{Ti}_{(\mathrm{s})}+2 \mathrm{MgCl}_{2(\mathrm{~s})}$
Calculate the percentage yield for this reaction.
29. When a sample of solid potassium chlorate is heated strongly, a decomposition reaction occurs. Solid potassium chloride and oxygen gas are produced.
(a) Write the balanced equation for this reaction.
(b) When this reaction was carried out, a mass of 3.78 g of potassium chloride remained after 7.62 g of potassium chlorate decomposed. Calculate the percentage yield of potassium chloride.
30. Sodium carbonate reacts with dilute hydrochloric acid, as shown by the following equation.
$\mathrm{Na}_{2} \mathrm{CO}_{3(\mathrm{aq})}+2 \mathrm{HCl}_{(\mathrm{aq})} \rightarrow 2 \mathrm{NaCl}_{(\mathrm{aq})}+\mathrm{H}_{2} \mathrm{O}_{(\ell)}+\mathrm{CO}_{2(\mathrm{~g})}$
(a) A chemist dissolves an impure sample of $\mathrm{Na}_{2} \mathrm{CO}_{3}$, with a mass of 0.250 g , in water. The chemist determines that 30.4 mL of 0.151 M HCl reacts with the $\mathrm{Na}_{2} \mathrm{CO}_{3}$ sample. Calculate the percentage purity of the sample.
(b) What volume of $\mathrm{CO}_{2}$ is produced, at $21.5^{\circ} \mathrm{C}$ and a pressure of 104.0 kPa , in the reaction described in part (a)?
31. When 15.0 g of copper and 4.83 g of sulfur are heated, a 13.7 g mass of copper(I) sulfide is produced.
$2 \mathrm{Cu}_{(\mathrm{s})}+\mathrm{S}_{(\mathrm{g})} \rightarrow \mathrm{Cu}_{2} \mathrm{~S}_{(\mathrm{s})}$
What is the percentage yield of $\mathrm{Cu}_{2} \mathrm{~S}$ ?
32. 130.4 mL of $0.459 \mathrm{~mol} / \mathrm{L}_{\mathrm{AgNO}}^{3}$ and 85.23 mL of $0.251 \mathrm{~mol} / \mathrm{L}^{\mathrm{AlCl}}{ }_{3}$ are mixed.
$3 \mathrm{AgNO}_{3(\text { aq) }}+\mathrm{AlCl}_{3 \text { (aq) }} \rightarrow \mathrm{Al}\left(\mathrm{NO}_{3}\right)_{3 \text { (aq) }}+3 \mathrm{AgCl}_{\text {(s) }}$
What mass of $\mathrm{AgCl}_{(\mathrm{s})}$ precipitates?
33. The following reaction occurs when a lead storage battery in a car is discharging.
$\mathrm{Pb}_{(\mathrm{s})}+\mathrm{PbO}_{2(\mathrm{~s})}+2 \mathrm{H}_{2} \mathrm{SO}_{4(\mathrm{aq})} \rightarrow 2 \mathrm{PbSO}_{4(\mathrm{~s})}+2 \mathrm{H}_{2} \mathrm{O}_{(\ell)}$
(a) A 3.850 g sample of $\mathrm{PbO}_{2}$ reacts completely with 2.710 mL of $\mathrm{H}_{2} \mathrm{SO}_{4}$. Calculate the concentration of $\mathrm{H}_{2} \mathrm{SO}_{4}$.
(b) What mass of $\mathrm{PbSO}_{4}$ is produced when 30.00 g of $\mathrm{H}_{2} \mathrm{SO}_{4}$ and 13.6 g of Pb react?
34. A sample of iron(III) oxide, with a mass of 325.0 g reacts with 90.75 L of carbon monoxide at $500.0^{\circ} \mathrm{C}$ and 1.216 atm .
$\mathrm{Fe}_{2} \mathrm{O}_{3(\mathrm{~s})}+\mathrm{CO}_{(\mathrm{g})} \rightarrow 2 \mathrm{Fe}_{(\mathrm{s})}+3 \mathrm{CO}_{2(\mathrm{~g})}$
(a) If a 185.0 g mass of iron is produced, what is the percentage yield for the reaction?
(b) What mass of reactant remains after the reaction stops?
35. The following reaction gives a $45.0 \%$ yield of manganese.
$2 \mathrm{Al}_{(\mathrm{s})}+3 \mathrm{MnO}_{(\mathrm{s})} \rightarrow \mathrm{Al}_{2} \mathrm{O}_{3(\mathrm{~s})}+3 \mathrm{Mn}_{(\mathrm{s})}$
What mass of $\mathrm{Mn}_{(\mathrm{s})}$ is produced when a 200.0 g sample of $\mathrm{Al}_{(\mathrm{s})}$ reacts with 300.0 g of $\mathrm{MnO}_{(\mathrm{s})}$ ?
36. What volume of $0.472 \mathrm{~mol} / \mathrm{L}_{\mathrm{AgNO}}^{3}$ will precipitate the chloride ion in 40.0 mL of $0.183 \mathrm{~mol} / \mathrm{L} \mathrm{AlCl}_{3}$ ?

## R. 14 Representing Organic Molecules

In organic chemistry, there are many different ways to represent the same molecule. You can represent an organic molecule using a molecular formula, an expanded molecular formula, or a structural diagram.

A molecular formula includes the actual number and type of atoms that are present in the molecule, but it gives no information about how the atoms are connected to each other. For example, the molecular formula, $\mathrm{C}_{2} \mathrm{H}_{6} \mathrm{O}$, indicates that there are 2 carbon atoms, 6 hydrogen atoms, and 1 oxygen atom in this molecule.

An expanded molecular formula shows the atoms in the order in which they appear in the molecule. You can write the molecular formula $\mathrm{C}_{2} \mathrm{H}_{6} \mathrm{O}$ in an expanded form as either $\mathrm{CH}_{3} \mathrm{CH}_{2} \mathrm{OH}$ or $\mathrm{CH}_{3} \mathrm{OCH}_{3}$. Each expanded form shows a different molecule that this molecular formula can represent. When writing an expanded molecular formula, use brackets to indicate groups that are attached to carbon atoms in the main chain. For example, the expanded molecular formula $\mathrm{CH}_{3} \mathrm{C}\left(\mathrm{CH}_{3}\right)_{2} \mathrm{CH}_{2} \mathrm{CH}_{3}$ represents a molecule with four carbon atoms attached in a chain, and two additional $-\mathrm{CH}_{3}$ groups attached to the second carbon atom.

A structural diagram is a simple drawing of a molecule. Structural diagrams include information about how the atoms are bonded. Figure R. 11 illustrates the three types of structural diagrams. Figure R. 12 shows structural diagrams for compounds with double and triple bonds.

- A complete structural diagram shows all the atoms in a structure and the way they are bonded to one another. Straight lines represent the bonds between the atoms.
- A condensed structural diagram is a more compact drawing of the structure. This type of diagram does not show the bonds to hydrogen atoms. Chemists assume that these bonds are present.
- A line structural diagram is even simpler than a condensed structural diagram. The end of each line, and the points at which the lines meet, represent carbon atoms. Hydrogen atoms are assumed to be present, but they are not shown. As you can see in Figure R.11(C) and Figure R.12, the lines that represent backbones of single-bonded or double-bonded carbon atoms are usually drawn in a zig-zag pattern. Triple-bonded carbon atoms are drawn in a straight line. (Note that lines are only used for the hydrocarbon portions of a molecule. Other atoms and groups, such as groups containing oxygen, nitrogen, and chlorine atoms, must be written in full.)



Figure R. 11 Complete (A), condensed (B), and line (C) structural diagrams for a compound with the molecular formula $\mathrm{C}_{5} \mathrm{H}_{12} \mathrm{O}$, and the expanded molecular formula $\mathrm{CH}_{3} \mathrm{CH}\left(\mathrm{CH}_{3}\right) \mathrm{CH}_{2} \mathrm{CH}_{2} \mathrm{OH}$



Figure R. 12 (A) a condensed structural diagram and the corresponding line structural diagram for an organic compound with the expanded molecular formula $\mathrm{CH}_{2}=\mathrm{C}\left(\mathrm{CH}_{3}\right) \mathrm{CH}_{2} \mathrm{CH}_{2} \mathrm{COOH}$; $(\mathrm{B})$ a condensed structural diagram and the corresponding line structural diagram for a compound with the expanded molecular formula $\mathbf{C H}_{3} \mathbf{C}=\mathbf{C C H}_{2} \mathbf{C H}_{3}$

## Practice Problems

37. What is the molecular formula of a compound with the expanded molecular formula $\mathrm{CH}_{3} \mathrm{CH}_{2} \mathrm{CH}(\mathrm{OH}) \mathrm{CH}_{2} \mathrm{CH}_{3}$ ?
38. Draw complete, condensed, and line structural diagrams for each compound.
(a) $\mathrm{CH}_{3} \mathrm{CH}_{2} \mathrm{CH}_{2} \mathrm{CH}_{2} \mathrm{CH}_{3}$
(c) $\mathrm{CH}_{2}=\mathrm{CHCH}_{2} \mathrm{CH}\left(\mathrm{CH}_{3}\right) \mathrm{CH}_{3}$
(b) $\mathrm{CH}_{3} \mathrm{CH}_{2} \mathrm{CH}\left(\mathrm{CH}_{3}\right) \mathrm{CH}_{2} \mathrm{CH}_{3}$
(d) $\mathrm{CH} \equiv \mathrm{CCH}_{2} \mathrm{CH}_{2} \mathrm{CH}_{3}$

## R. 15 Isomers of Organic Compounds

Many molecular formulas represent more than one molecular structure. Compounds that have the same molecular formula, but different structures, are called structural isomers. As you saw earlier, the molecular formula $\mathrm{C}_{2} \mathrm{H}_{6} \mathrm{O}$ can be represented by the expanded molecular formulas $\mathrm{CH}_{3} \mathrm{CH}_{2} \mathrm{OH}$ and $\mathrm{CH}_{3} \mathrm{OCH}_{3}$. The atoms in these molecules are attached differently to form two different structures. Both molecules, however, have the same molecular formula. Thus, they are isomers of each other. Because isomers have different shapes and bonding, they usually have different physical and chemical properties.

A molecule cannot rotate around a $\mathrm{C}=\mathrm{C}$ bond. This fact makes a different type of isomer possible. Cis-trans isomers (also called geometric isomers) are compounds that have the same molecular formula, but different arrangements of atoms around a double carbon-carbon bond.

- In a cis-isomer, the two largest groups are on the same side of the double bond.
- In a trans-isomer, the two largest groups are on different sides of the double bond.

Figure R.13, in the margin, illustrates a pair of geometric isomers.


Figure R. 13 Geometric isomers of the compound 2-butene

## Practice Problems

39. Identify each diagram as a cis-isomer or a trans-isomer.
(a)

(b)

(c)

40. (a) Draw condensed structural diagrams for five isomers with the molecular formula $\mathrm{C}_{6} \mathrm{H}_{12}$.
(b) Draw line structural diagrams for five new isomers that also have the molecular formula $\mathrm{C}_{6} \mathrm{H}_{12}$. Include one pair of cis-trans isomers.

## Answers to Practice Problems

1.(a) xenon tetrafluoride (b) phosphorus pentachloride (phosphorus(V) chloride) (c) carbon monoxide (d) sulfur hexafluoride (e) dinitrogen tetroxide (f) sodium fluoride (g) calcium carbonate (h) barium thiosulfate (i) nitrogen triiodide 2.(a) titanic oxide; titanium(IV) oxide (b) cobaltous chloride; cobalt(II) chloride (c) nickelous bromide; nickel(II) bromide (d) mercuric oxide; mercury(II) oxide (e) stannous chloride; tin(II) chloride 3.(a) $\mathrm{Cu}(\mathrm{OH})_{2}$ (b) $\mathrm{Hg}_{3} \mathrm{~N}_{2}$ (c) $\mathrm{Mn}\left(\mathrm{HCO}_{3}\right)_{2}$ (d) $\mathrm{SBr}_{6}$ (e) $\mathrm{K}_{2} \mathrm{CrO}_{4}$ (f) $\mathrm{AsCl}_{3}$ (g) $\mathrm{Fe}\left(\mathrm{C}_{2} \mathrm{H}_{3} \mathrm{O}_{2}\right)_{3}$ (h) $\mathrm{K}_{2} \mathrm{O}_{2}$ (i) $\left(\mathrm{NH}_{4}\right)_{2} \mathrm{Cr}_{2} \mathrm{O}_{7}$ (j) $\mathrm{Sn}\left(\mathrm{MnO}_{4}\right)_{2}$ (k) $\mathrm{Pb}\left(\mathrm{NO}_{3}\right)_{2}$. (a) $\mathrm{Zn}\left(\mathrm{ClO}_{4}\right)_{2}$ (b) $\mathrm{Hg}_{3} \mathrm{~N}_{2}$ (c) $\mathrm{SnF}_{2}$ (d) $\mathrm{NH}_{4} \mathrm{H}_{2} \mathrm{PO}_{3}$ (e) $\mathrm{Mn}\left(\mathrm{SiO}_{4}\right)_{2}$ (f) $\mathrm{Fe}(\mathrm{OH})_{3}$ (g) CoN (h) $\mathrm{Pb}\left(\mathrm{HCO}_{3}\right)_{2}$ (i) $\mathrm{Ag}_{2} \mathrm{C}_{2} \mathrm{O}_{4}$ (j) $\mathrm{PtO}_{2}$ (k) SiC (I) $\mathrm{Ni}_{2}\left(\mathrm{SO}_{3}\right)_{3}$ ( $\mathbf{m}$ ) $\mathrm{SnCO}_{3}$ (n) $\mathrm{Al}\left(\mathrm{MnO}_{4}\right)_{3}$ 5.(a) $\mathrm{H}_{2} \mathrm{O}_{(\ell)}+3 \mathrm{NO}_{2(\mathrm{~g})} \rightarrow 2 \mathrm{HNO}_{3(\mathrm{aq})}+\mathrm{NO}_{(\mathrm{g})}$ (b) $2 \mathrm{NaOH}_{(\text {aq) }}+\mathrm{Cl}_{2(\mathrm{~g})} \rightarrow \mathrm{NaCl}_{(\text {aq })}+\mathrm{NaClO}_{(\text {aq) }}+\mathrm{H}_{2} \mathrm{O}_{(\ell)}$ (c) $2\left(\mathrm{NH}_{4}\right)_{3} \mathrm{PO}_{4(\mathrm{aq})}+3 \mathrm{Ba}(\mathrm{OH})_{2(\mathrm{aq})} \rightarrow \mathrm{Ba}_{3}\left(\mathrm{PO}_{4}\right)_{2(\mathrm{~s})}+6 \mathrm{NH}_{4} \mathrm{OH}_{\text {(aq) }}$ (d) $2 \mathrm{Li}_{(\mathrm{s})}+2 \mathrm{H}_{2} \mathrm{O}_{(\ell)} \rightarrow 2 \mathrm{LiOH}_{(\mathrm{aq})}+\mathrm{H}_{2(\mathrm{~g})}(\mathrm{e}) 4 \mathrm{NH}_{3(\mathrm{~g})}+3 \mathrm{O}_{2(\mathrm{~g})} \rightarrow 2 \mathrm{~N}_{2(\mathrm{~g})}+6 \mathrm{H}_{2} \mathrm{O}_{(\ell)}$ (f) $\mathrm{Sb}_{4} \mathrm{~S}_{6(\mathrm{~s})}+9 \mathrm{O}_{2(\mathrm{~g})} \rightarrow \mathrm{Sb}_{4} \mathrm{O}_{6(\mathrm{~s})}+6 \mathrm{SO}_{2(\mathrm{~g})}$ (g) $2 \mathrm{Al}\left(\mathrm{NO}_{3}\right)_{3(\text { aq) }}+3 \mathrm{H}_{2} \mathrm{SO}_{4(\text { aq })} \rightarrow \mathrm{Al}_{2}\left(\mathrm{SO}_{4}\right)_{3(\mathrm{aq})}+6 \mathrm{HNO}_{3(\text { aq })}$ (h) $2 \mathrm{Cu}\left(\mathrm{NO}_{3}\right)_{2(\mathrm{~s})} \rightarrow 2 \mathrm{CuO}_{(\mathrm{s})}+4 \mathrm{NO}_{2(\mathrm{~g})}+\mathrm{O}_{2(\mathrm{~g})}$ (i) $2 \mathrm{KI}_{\text {(aq) }}+\mathrm{MnO}_{2(\mathrm{~s})}+2 \mathrm{H}_{2} \mathrm{SO}_{4 \text { (aq) }} \rightarrow \mathrm{K}_{2} \mathrm{SO}_{4 \text { (aq) }}+\mathrm{MnSO}_{4(\text { aq) }}+2 \mathrm{H}_{2} \mathrm{O}_{(\ell)}+\mathrm{I}_{2(\mathrm{~s})}$ (j) $2 \mathrm{C}_{6} \mathrm{H}_{6(\ell)}+15 \mathrm{O}_{2(\mathrm{~g})} \rightarrow 12 \mathrm{CO}_{2(\mathrm{~g})}+6 \mathrm{H}_{2} \mathrm{O}_{(\mathrm{g})}$. (a) decomposition (b) synthesis (c) single displacement 7.(a) $\mathrm{MgSO}_{3(\mathrm{~s})} \rightarrow \mathrm{Mg}^{2+}{ }_{\text {(aq) }}+\mathrm{SO}_{3}{ }^{2-}{ }_{\text {(aq) }}$
(b) $\mathrm{Fe}(\mathrm{OH})_{2(\mathrm{~s})} \rightarrow \mathrm{Fe}^{2+}{ }_{\text {(aq) }}+2 \mathrm{OH}^{-}{ }_{\text {(aq) }}$ (c) $\mathrm{Sn}\left(\mathrm{SO}_{4}\right)_{2(\mathrm{~s})} \rightarrow \mathrm{Sn}^{4+}{ }_{\text {(aq) }}+2 \mathrm{SO}_{4}{ }^{2-}{ }_{(\text {aq) }}$
(d) $\mathrm{Ag}_{2} \mathrm{~S}_{\text {(s) }} \rightarrow 2 \mathrm{Ag}^{+}{ }_{\text {(aq) }}+\mathrm{S}^{2-}{ }_{\text {(aq) }}$ (e) $\mathrm{Al}_{2}\left(\mathrm{C}_{2} \mathrm{O}_{4}\right)_{3(\mathrm{~s})} \rightarrow 2 \mathrm{Al}^{3+}{ }_{\text {(aq) }}+3 \mathrm{C}_{2} \mathrm{O}_{4}{ }^{2-}{ }_{\text {(aq) }}$
8.(a) $\mathrm{Pb}\left(\mathrm{NO}_{3}\right)_{2(\text { s })} \rightarrow \mathrm{Pb}^{2+}{ }_{\text {(aq) }}+2 \mathrm{NO}_{3}^{-}{ }_{\text {(aq) }} ; \mathrm{ZnI}_{2 \text { (s) }} \rightarrow \mathrm{Zn}^{2+}{ }_{\text {(aq) }}+2 \mathrm{I}^{-}{ }_{\text {(aq) }}$
$\mathrm{Pb}^{2+}{ }_{\text {(aq) }}+2 \mathrm{NO}_{3}^{-}{ }_{\text {(aq) }}+\mathrm{Zn}^{2+}{ }_{\text {(aq) }}+2 \mathrm{I}^{-}{ }_{\text {(aq) }} \rightarrow \mathrm{PbI}_{2(\mathrm{~s})}+\mathrm{Zn}^{2+}{ }_{(\text {aq })}+2 \mathrm{NO}_{3}^{-}{ }_{\text {(aq) }}$
$\mathrm{Pb}^{2+}{ }_{\text {(aq) }}+2 \mathrm{I}^{-}{ }_{\text {(aq) }} \rightarrow \mathrm{PbI}_{2(\mathrm{~s})}$; spectator ions: $\mathrm{Zn}^{2+}{ }_{\text {(aq) }}$ and $\mathrm{NO}_{3}{ }^{-}{ }_{\text {(aq) }}$
(b) $\mathrm{KClO}_{3(\mathrm{~s})} \rightarrow \mathrm{K}^{+}{ }_{\text {(aq) }}+\mathrm{ClO}_{3}^{-}{ }_{(\text {aq) }} ; \mathrm{CaCl}_{2(\mathrm{~s})} \rightarrow \mathrm{Ca}^{2+}{ }_{\text {(aq) }}+2 \mathrm{Cl}^{-}{ }_{(\text {aq })}$
$\mathrm{K}^{+}{ }_{\text {aq) }}+\mathrm{ClO}_{3}^{-}{ }_{(\text {aq) }}+\mathrm{Ca}^{2+}{ }_{\text {(aq) }}+2 \mathrm{Cl}^{-}{ }_{\text {(aq) }} \rightarrow \mathrm{NR}$
(c) $\left(\mathrm{NH}_{4}\right)_{3} \mathrm{PO}_{4(\mathrm{~s})} \rightarrow 3 \mathrm{NH}_{4}{ }^{+}$(aq) $+\mathrm{PO}_{4}{ }^{3-}{ }_{\text {(aq) }} ; \mathrm{Cu}\left(\mathrm{ClO}_{3}\right)_{2(\mathrm{~s})} \rightarrow \mathrm{Cu}^{2+}{ }_{\text {(aq) }}+2 \mathrm{ClO}_{3}{ }^{-}{ }_{\text {(aq) }}$ $6 \mathrm{NH}_{4}{ }^{+}{ }_{(\mathrm{aq})}+2 \mathrm{PO}_{4}{ }^{3-}{ }_{(\mathrm{aq})}+3 \mathrm{Cu}^{2+}{ }_{(\mathrm{aq})}+6 \mathrm{ClO}_{3}{ }^{-}{ }_{\text {(aq) }} \rightarrow \mathrm{Cu}_{3}\left(\mathrm{PO}_{4}\right)_{2(\mathrm{~s})}+6 \mathrm{NH}_{4}{ }^{+}$(aq) $+6 \mathrm{ClO}_{3}{ }^{-}{ }_{\text {(aq) }}$; $3 \mathrm{Cu}^{2+}{ }_{\text {(aq) }}+2 \mathrm{PO}_{4}{ }^{3-}{ }_{(\text {aq })} \rightarrow \mathrm{Cu}_{3}\left(\mathrm{PO}_{4}\right)_{2(\mathrm{~s})}$; spectator ions: $\mathrm{NH}_{4}{ }^{+}{ }_{(\text {aq })}$ and $\mathrm{ClO}_{3}{ }^{-}{ }^{(\text {aq })}$ (d) $\mathrm{NaOH}_{(\mathrm{s})} \rightarrow \mathrm{Na}^{+}{ }_{(\text {aq })}+\mathrm{OH}^{-}{ }_{(\text {aq })} ; \mathrm{Fe}\left(\mathrm{NO}_{3}\right)_{3(\mathrm{~s})} \rightarrow \mathrm{Fe}^{3+}{ }_{(\text {aq })}+3 \mathrm{NO}_{3}^{-}{ }_{(\text {aq })}$ $3 \mathrm{Na}^{+}{ }_{\text {(aq) }}+3 \mathrm{OH}^{-}{ }_{(\text {aq })}+\mathrm{Fe}^{3+}{ }_{\text {(aq) }}+3 \mathrm{NO}_{3}{ }^{-}{ }_{\text {(aq) }} \rightarrow \mathrm{Fe}(\mathrm{OH})_{3(\mathrm{~s})}+3 \mathrm{Na}^{+}{ }_{\text {(aq) }}+3 \mathrm{NO}_{3}{ }^{-}{ }_{\text {(aq) }}$ $\mathrm{Fe}^{3+}{ }_{(\mathrm{aq})}+3 \mathrm{OH}^{-}{ }_{(\text {aq) }} \rightarrow \mathrm{Fe}(\mathrm{OH})_{3(\mathrm{~s})} ;$ spectator ions: $\mathrm{Na}^{+}{ }_{(\text {aq) }}$ and $\mathrm{NO}_{3}{ }^{-}{ }_{(\text {aq) }}$ (e) $\mathrm{Ca}\left(\mathrm{NO}_{3}\right)_{2(\mathrm{~s})} \rightarrow \mathrm{Ca}^{2+}{ }_{\text {(aq) }}+2 \mathrm{NO}_{3}^{-}{ }_{(\text {aq) }} ; \mathrm{K}_{2} \mathrm{CO}_{3(\mathrm{~s})} \rightarrow 2 \mathrm{~K}^{+}{ }_{\text {(aq) }}+\mathrm{CO}_{3}{ }^{2-}{ }_{\text {(qq) }}$ $\mathrm{Ca}^{2+}{ }_{\text {(qq) }}+2 \mathrm{NO}_{3}{ }^{-}{ }_{\text {(qq) }}+2 \mathrm{~K}^{+}{ }_{\text {(qq) }}+\mathrm{CO}_{3}{ }^{2-}{ }_{\text {(aq) }} \rightarrow \mathrm{CaCO}_{3(\mathrm{~s})}+2 \mathrm{~K}^{+}{ }_{\text {(aq) }}+2 \mathrm{NO}_{3}{ }^{-}$(aq) $\mathrm{Ca}^{2+}{ }_{(\text {aq })}+\mathrm{CO}_{3}{ }^{2-}{ }_{(\text {aq })} \rightarrow \mathrm{CaCO}_{3(\mathrm{~s})}$; spectator ions: $\mathrm{K}^{+}{ }_{(\text {aq })}$ and $\mathrm{NO}_{3}{ }^{-}$(aq)
9. $2.55 \times 10^{-2} \mathrm{~g}$ 10. 5.84 mol 11. $1.6 \times 10^{-3} \mathrm{~mol}$ 12. 20.0 g 13. $2.68 \times 10^{22}$ atoms
14. $1.98 \times 10^{-4} \mathrm{~g}$ 15. sample A: $3.39 \times 10^{22}$ atoms; sample B: $3.38 \times 10^{22}$ atoms
16. $4.38 \times 10^{3} \mathrm{~mol}$ 17. 1.31 L 18. $0.360 \mathrm{~mol} / \mathrm{L} 19.0 .71 \mathrm{~g}$ 20. $0.299 \mathrm{~L} 21.1 .66 \times 10^{-2} \mathrm{~g}$
22. $1.82 \mathrm{~mol} / \mathrm{L}$ 23. $0.470 \mathrm{~mol} / \mathrm{L} 24$. For equal masses of solute, volume and mass of the solution are the same. 25.(a) $4.49 \times 10^{-5} \mathrm{~mol} / \mathrm{L}$ (b) $9.31 \mathrm{ppm} \mathrm{Pb}^{2+}$
(c) $1.49 \times 10^{-3} \mathrm{~m} / \mathrm{m}$ percent 26.32 .17 g 27. $0.332 \mathrm{~g} 28.94 .43 \%$
29.(a) $2 \mathrm{KClO}_{3(\mathrm{~s})} \rightarrow 2 \mathrm{KCl}_{(\mathrm{s})}+3 \mathrm{O}_{2(\mathrm{~g})}$ (b) $81.5 \% \quad 30$. (a) $97.6 \%$ (b) $54.2 \mathrm{~mL} 31.72 .9 \% \quad 32.8 .57 \mathrm{~g}$
33.(a) $11.88 \mathrm{~mol} / \mathrm{L}$ (b) 39.8 g 34.(a) $95.26 \%$ (b) 47.30 g of $\mathrm{Fe}_{2} \mathrm{O}_{3} \quad 35.105 \mathrm{~g} 36.46 .6 \mathrm{~mL}$
37. $\mathrm{C}_{5} \mathrm{H}_{12} \mathrm{O}$
38.(a)

$\mathrm{CH}_{3}-\mathrm{CH}_{2}-\mathrm{CH}_{2}-\mathrm{CH}_{2}-\mathrm{CH}_{3}$




(c)



$\mathrm{CH} \equiv \mathrm{C}-\mathrm{CH}_{2}-\mathrm{CH}_{2}-\mathrm{CH}_{3}$

39.(a) cis (b) trans (c) trans
40.(a) $\mathrm{CH}_{2}=\mathrm{CH}-\mathrm{CH}_{2}-\mathrm{CH}_{2}-\mathrm{CH}_{2}-\mathrm{CH}_{3}$,
$\mathrm{CH}_{3}-\mathrm{CH}=\mathrm{CH}-\mathrm{CH}_{2}-\mathrm{CH}_{2}-\mathrm{CH}_{3}$,
$\mathrm{CH}_{3}-\mathrm{CH}-\mathrm{CH}=\mathrm{CH}-\mathrm{CH}_{2}-\mathrm{CH}_{3}$,








[^0]:    Figure R. 6 Lewis structures showing the formation of an ionic bond between calcium and fluorine

