When you buy a carton that says it contains one dozen eggs, you know that the carton contains 12 eggs. The term “dozen” is a counting unit. You use counting units for convenience all the time. For example, you group your socks into counting units of two, called pairs. You might buy a six-pack of soft drinks.

### The Mole

For convenience, chemists group particles by large numbers that are easier to work with. The numbers of elementary entities (atoms, molecules, or formula units) in even a small sample of a substance are enormous. Therefore, they are grouped by an extremely large number. The grouping that chemists use is called the mole.

- The mole (symbol mol) is defined as the amount of substance that contains as many elementary entities (atoms, molecules, or formula units) as exactly 12 g of carbon-12, the most common isotope of carbon.
- One mole (1 mol) of a substance has been determined to contain $6.022 \times 10^{23}$ elementary entities of the substance. The number $6.022 \times 10^{23}$ is called Avogadro’s number.

Notice that there are nine significant digits in the most precise and accurate determination of Avogadro’s number to date. You will rarely need this level of precision. In this textbook, you will use the value $6.02 \times 10^{23}$. For example, there are $6.02 \times 10^{23}$ carbon atoms in 1 mol of carbon, C. There are $6.02 \times 10^{23}$ carbon dioxide molecules in 1 mol of carbon dioxide, CO$_2$. There are $6.02 \times 10^{23}$ sodium chloride formula units in 1 mol of sodium chloride, NaCl(s).

### The Magnitude of Avogadro’s Number

The number $6.02 \times 10^{23}$ is enormous. One way to appreciate the vast size of the number is to relate it to everyday objects that you can visualize. Figures 3.22 and 3.23 provide some examples. In Think and Link Investigation 3-E, you will visualize covering the surface of Alberta with $6.02 \times 10^{23}$ pennies.
How Big Is $6.02 \times 10^{23}$?

Think About It

In this investigation, you will consider the following scenario. You have $6.02 \times 10^{23}$ pennies, which you are going to use to cover the surface of Alberta. What will be the height of the layer of pennies? You will stack the pennies in columns of roughly equal height. The columns must touch at the bottom.

What to Do

1. As a group, design a procedure for answering the question above. You will need to consider several variables, including the following:
   - How will you determine the surface area of Alberta?
   - How will you lay out the pennies?
   - How will you determine the height of a penny?

2. Carry out your procedure to determine the height of the layer of $6.02 \times 10^{23}$ pennies.

3. Write a detailed report showing what you did and how you arrived at your answer. Pay careful attention to significant digits in your measurements and calculations. Include answers to the Analyze questions below in your report.

Analyze

1. To answer the question, you probably needed to make one or more approximations. List the approximations you made. How do you think these approximations affected your estimate (that is, did they increase or decrease your final answer)?

2. Compare your group’s procedure and results with the procedure and results of other groups. Did the answers differ? By how much did they differ? How did the procedures differ? Which group do you think came up with the best procedure, and why?

3. How could you have improved your procedure for this investigation?

Extend Your Knowledge and Skills

4. With your group, create your own problem using Avogadro’s number and challenge other groups to solve it. Here are examples to give you some ideas:
   - If you have $6.02 \times 10^{23}$, and you share it equally among all the people in the world, how much money would each person receive?
   - How does the mass of $6.02 \times 10^{23}$ bananas compare with the mass of Earth?
   - If you removed $6.02 \times 10^{23}$ mL of seawater from the world’s oceans, would the oceans be completely emptied?

5. Choose and solve one of the three problems given as examples above.
Relating Mass and Amount

On its own, thinking of atoms and molecules in amounts of $6.02 \times 10^{23}$ particles is of little practical use. There needs to be a convenient way to measure amounts of substances in moles. Fortunately, it is possible to calculate what mass of a substance is equal to one mole of that substance. To do this, chemists use the relative masses of elements.

![INTERNET CONNECT](www.mcgrawhill.ca/links/sciencefocus10)

The mole has its own day! Mole Day begins at 6:02 A.M. and ends at 6:02 P.M. on October 23 each year. Why are those times and that date appropriate? High school teachers and students throughout the world take part by writing jokes, songs, and poems, and by participating in projects to celebrate the mole and chemistry in general. What is the theme of this year’s Mole Day and how can you participate? Go to the web site above to find out where to go next. Prepare a brief proposal outlining three possible ways your class could get involved in Mole Day.

Can You Count on It?

Chemists and chemical engineers usually need to control the number of atoms, molecules, and ions in their reactions very carefully. Chemists can calculate the approximate numbers of the particles by measuring their masses. How accurate is this method?

**Materials**
- small identical items to be counted (e.g., paper clips or bingo chips)
- re-sealable plastic bag
- electronic balance

**Procedure**

1. Count out a small number of items and measure their mass. Record the number of items and the mass.
2. Determine and record the mass of an empty re-sealable plastic bag.
3. Fill the bag to the brim with items and seal it.
4. Determine and record the mass of the filled plastic bag.
5. Develop and carry out a method to determine the number of items in the bag without opening it.

**What Did You Find Out?**

1. According to your determination, how many items are in the bag?
2. Explain how you determined that number.
3. After calculating the number of items in the bag, open the bag and count the number of items. Compare your calculated number of items with the actual number of items.
4. Discuss the accuracy of your method of determining the number of items in the plastic bag. What are some sources of error that may have affected your results?
5. In this activity, you were able to count manually a group of small items and measure the total mass of the group to determine an average mass. Do you think chemists are able to determine the average mass of an atom or molecule in the same way? Explain your answer.
The Atomic Molar Mass of an Element

The atomic molar mass ($M$) of an element is a weighted average of the mass of 1 mol of all of the naturally occurring isotopes of the element. You can find the atomic molar mass listed for each element on the periodic table. For example, the atomic molar mass of iron is listed on the periodic table as 55.85 g/mol. In other words, the mass of 1 mol of iron is 55.85 g. Similarly, the atomic molar mass of sodium is 22.99 g/mol. Therefore, the mass of 1 mol of sodium is 22.99 g. Figure 3.24 shows samples of elements. Each sample contains $6.02 \times 10^{23}$ atoms, or 1 mol.

![Figure 3.24](image)

Each sample above contains $6.02 \times 10^{23}$ atoms. Note that the mass of each sample is different.

Some elements exist as molecules, not atoms. For example, the element nitrogen exists as a molecule composed of two nitrogen atoms, $N_2(g)$. Therefore, 1 mol of nitrogen molecules contains 2 mol of nitrogen atoms. The molar mass of nitrogen molecules is therefore twice the atomic molar mass of nitrogen as shown on the periodic table: $2 \times 14.01$ g/mol = 28.02 g/mol. In other words, 1 mol of molecular nitrogen has a mass of 28.02 g.

The Molar Mass of a Compound

The term molar mass ($M$), with units of g/mol, is used to refer to the mass of 1 mol of any pure substance. You can determine the molar mass of a compound by using the formula of the compound. For example, the formula of carbon dioxide, $CO_2(g)$, tells you that each molecule contains one carbon atom and two oxygen atoms.
By extension, 1000 molecules of CO$_2$ contain 1000 carbon atoms and $2 \times 1000$ oxygen atoms. Further, $6.02 \times 10^{23}$ molecules of CO$_2$ contain $6.02 \times 10^{23}$ carbon atoms and $2 \times 6.02 \times 10^{23}$ oxygen atoms. In other words, 1 mol of CO$_2$ contains 1 mol of carbon atoms and 2 mol of oxygen atoms. Therefore, the molar mass of a molecular compound may be calculated as follows:

mass of 1 mol of CO$_2$ = (1 mol C $\times$ M$_C$) + (2 mol O $\times$ M$_O$) = (1 mol C $\times$ 12.01 g/mol C) + (2 mol O $\times$ 16.00 g/mol O) = 44.01 g

Therefore, the molar mass of carbon dioxide is 44.01 g/mol.

You can calculate the molar mass of an ionic compound in the same way:

mass of 1 mol of Mg(NO$_3$)$_2$ = (1 mol Mg $\times$ M$_{Mg}$) + (2 mol N $\times$ M$_N$) + (6 mol O $\times$ M$_O$) = (1 mol Mg $\times$ 24.31 g/mol Mg) + (2 mol N $\times$ 14.01 g/mol N) + (6 mol O $\times$ 16.00 g/mol O) = 148.33 g

Therefore, the molar mass of magnesium nitrate is 148.33 g/mol.

Figure 3.25 shows some examples of molar amounts of substances and their masses. Try the following problems to practise determining molar masses.

**Practice Problems**

9. Find the molar mass of each of the following elements.
   (a) potassium, K$_{(s)}$
   (b) zirconium, Zr$_{(s)}$
   (c) chlorine, Cl$_2$(g)
   (d) oxygen, O$_2$(g)

10. Determine the molar mass of each of the following compounds.
    (a) potassium bromide, KBr$_{(s)}$
    (b) methane, CH$_4$(g)
    (c) sodium sulfate, Na$_2$SO$_4$(s)
    (d) aluminium nitrate, Al(NO$_3$)$_3$(s)
Converting Between Mass and Moles

It is often useful for chemists to convert between the amount of a sample expressed in moles and the mass of the sample expressed in grams. To do this, you can use a method called the **factor label method**. The factor label method uses conversion factors to change units without affecting the value. You can use conversion factors to change an amount of a sample expressed in moles to the mass of the sample, expressed in grams. In this case, the conversion factor is the molar mass of the substance. For example, to determine the mass of 2 mol of helium, use the molar mass of helium, 4.00 g/mol, as a conversion factor. You can do this because 4.00 g of helium is equivalent to 1 mol of helium.

conversion factor = \( \frac{4.00 \text{ g He}}{1 \text{ mol He}} \)

Multiply 2 mol by the conversion factor to determine its mass.

mass of helium = \( 2 \text{ mol He} \times \frac{4.00 \text{ g He}}{1 \text{ mol He}} \)

= 8.00 g He

Notice that the answer has the desired units, grams. Model Problem 1 shows how to use this method to convert from amount to mass of a compound. You can use the inverse of the molar mass to convert from mass to amount, as shown in Model Problem 2.

Alternatively, you may use the following equation to convert between mass and moles:

\[ n = \frac{m}{M} \]

\( n \) = amount (mol)
\( m \) = mass (g)
\( M \) = molar mass (g/mol)

---

**Model Problem 1**

What is the mass in grams of 7.50 mol of \( \text{H}_2\text{O}(l) \)?

**Solution**

number of moles of \( \text{H}_2\text{O}(l) = 7.50 \text{ mol} \)

amount of water (in mol) \( \rightarrow \) mass of water (in g)

Use molar mass as a conversion factor to convert from moles to grams. First, calculate the molar mass of water.

mass of 1 mol of \( \text{H}_2\text{O}(l) \)

\[ = 2 \text{ mol H} \times 1.01 \frac{\text{g}}{\text{mol H}} + 1 \text{ mol O} \times 16.00 \frac{\text{g}}{\text{mol O}} \]

= 18.02 g

Therefore, the molar mass of water is 18.02 g/mol.

Set up an equation, using molar mass as a conversion factor. Check that the units cancel, leaving the desired unit (grams).

\[ 7.50 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g}}{1 \text{ mol H}_2\text{O}} = 1.35 \times 10^2 \text{ g} \]

The mass of 7.50 mol of water is \( 1.35 \times 10^2 \text{ g} \).
Most people are familiar with stories of Olympic™ athletes who enjoy the triumph after being declared the winner in their sport. The triumph is short-lived for a few athletes, however, if the gold medal is denied them because they have tested positive for performance-enhancing drugs. Who conducts the tests for these substances, and how are these tests conducted?

Dr. Christiane Ayotte has been the head of Canada’s Doping Control Laboratory since 1991. When a urine sample arrives at the doping control lab, Dr. Ayotte and her team ensure careful handling of the sample. Portions of it are taken for six different analytical procedures. The more than 150 substances banned by the international Olympic Committee (IOC) are grouped according to their physical and chemical properties. There are two main steps in analyzing a sample:

1. purification, which involves steps such as filtration and extraction using solvents, and
2. analysis by gas chromatography, mass spectrometry, or high-performance liquid chromatography. Chromatography is a part of the processes whereby chemists separate mixtures into pure substances.

Just the presence of most banned substances in a urine sample means a positive result. Other substances must be present in an amount higher than a certain threshold. This is when knowing about chemical amounts comes in handy. A male athlete would have to consume 10 very strong cups of coffee within 15 min to go over the limit for caffeine. Ephedrines and pseudoephedrines, two decongestants that are found in cough remedies and act as stimulants, have a cut-off level that allows athletes to take them up to one or two days before a competition.

Dr. Ayotte and her team are constantly searching for reliable tests for natural substances, developing new analytical techniques, and determining the normal levels of banned substances for male and female athletes. These are just part of the challenges she and her team face. Dr. Ayotte must also defend her tests in hearings and with the press, especially when high-profile athletes are involved. Dr. Ayotte’s findings are sometimes unpleasant to report. However, she believes that integrity and a logical mind are essential aspects of being a good scientist.

Dr. Christiane Ayotte

Model Problem 2

The mass of an iron bar is 16.8 g. What amount (in mol) of iron is in the sample?

Solution

mass of Fe = 16.8 g
mass of iron (in g) → amount of iron (in mol)

Set up an equation to convert from mass to moles, using the molar mass of iron in your conversion factor. Check that the answer has the desired unit (moles).

\[
16.8 \text{ g Fe} \times \frac{1 \text{ mol}}{55.85 \text{ g Fe}} = 0.301 \text{ mol}
\]

A 16.8 g sample of iron contains 0.301 mol of iron.

Practice Problems

11. Determine the mass of 43.2 mol magnesium, Mg(s).
12. What is the mass of \(9.01 \times 10^3\) mol carbon tetrafluoride, CF\(_4\)(g)?
13. How many moles are in a 0.021 g sample of helium, He(g)?
14. Determine how many moles of potassium chlorate, KClO\(_3\)(s), are in a 344.9 g sample of the pure substance.
Working with the Mole

The best way to get better at working with the mole is to practise. Sharpen your skills by working through the following Practice Problems.

**Practice Problems**

15. Determine the molar mass of the following elements.
   (a) cesium, Cs(t')
   (b) gold, Au(s)
   (c) hydrogen, H2(g)
   (d) nitrogen, N2(g)

16. Determine the molar mass of each of the following compounds.
   (a) NO3(g)
   (b) Al2S3(s)
   (c) FeSO4(s)
   (d) (NH4)2CO3(s)

17. What is the mass, in grams, of 8.52 mol of each of the following substances?
   (a) H2O(t)
   (b) Al(s)
   (c) NiCl2(s)
   (d) O2(g)

18. What is the amount, in moles, of 342.5 g of each of the following substances?
   (a) W(s)
   (b) NCl3(t)
   (c) S8(s)
   (d) Ca(NO3)2(s)

19. Express the following as molar amounts.
   (a) 10.00 g CO2(g)
   (b) 3.50 g C5H12(g)
   (c) 14.00 g Cr(OH)3(s)
   (d) 6.97 g Al(NO3)3(s)

20. Re-order each of the following groups of substances from lowest to highest molar mass.
   (a) Nb(s), Tl(s), U(s), Sb(s)
   (b) F2(g), I2(g), Cl2(g), Br2(t)
   (c) Ir(s), Se(s), GaP(s), HClO3(s)
   (d) NH3(g), SO2(g), K3PO4(s), Pb(s)

21. A student writes down the following information.
    mass of beaker = 52.43 g
    mass of beaker and CuSO4(s) = 65.41 g
    What amount of CuSO4(s), in moles, did the student add to the beaker?

22. A student uses a graduated cylinder to measure 25.5 mL of water at 25°C. The density of water is 1.00 g/mL at 25°C.
   (a) What is the mass of the water in the graduated cylinder?
   (b) What amount of water, in moles, is in the graduated cylinder?
The Mole and the Law of Conservation of Mass

Using the mole concept, you can relate the coefficients in balanced chemical equations to the mass of the substances involved. Start by thinking about how the coefficients relate to moles of substances. For example, consider this chemical equation:

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

You can multiply all of the coefficients by two and the equation will remain balanced. Four molecules of hydrogen react with two molecules of oxygen to form four molecules of water. There are then eight hydrogen atoms on the reactant side and on the product side. There are four oxygen atoms on the reactant side and on the product side. In fact, as shown below, you can multiply the coefficients by any factor you like.

\[
\begin{array}{ccc}
2\text{H}_2(g) & + & \text{O}_2(g) & \rightarrow & 2\text{H}_2\text{O}(l) \\
2 \text{ molecules} & 1 \text{ molecule} & 2 \text{ molecules} \\
2 \times 2 \text{ molecules} & 1 \times 2 \text{ molecules} & 2 \times 2 \text{ molecules} \\
2 \times 1000 \text{ molecules} & 1 \times 1000 \text{ molecules} & 2 \times 1000 \text{ molecules} \\
2 \times 6.02 \times 10^{23} \text{ molecules} & 1 \times 6.02 \times 10^{23} \text{ molecules} & 2 \times 6.02 \times 10^{23} \text{ molecules} \\
2 \text{ mol} & 1 \text{ mol} & 2 \text{ mol}
\end{array}
\]

Notice that \(2 \times 6.02 \times 10^{23}\) molecules is the same as 2 mol of molecules. In the same way, \(1 \times 6.02 \times 10^{23}\) molecules is the same as 1 mol of molecules.

Now you have two ways to interpret the coefficients of a balanced chemical equation. The coefficients can represent the number of particles (molecules or formula units) involved in the reaction. Or, the same coefficients can represent the amount in moles of each substance involved in the reaction.

In Table 3.5, you can see how molecules, moles, and mass are related in the chemical equation. Notice that the total mass of the product is equal to the total mass of the reactants. If you carry out this comparison with any balanced chemical equation, you will see the same thing. The total mass of the substances on the products side equals the total mass of the substances on the reactants side. You know from Section 3.3 that balanced chemical equations obey the law of conservation of mass when the coefficients refer to atoms, molecules, or formula units. Table 3.5 shows that balanced chemical equations also obey the law of conservation of mass when they refer to moles.

Table 3.5 Chemical Equations and The Law of Conservation of Mass

<table>
<thead>
<tr>
<th>Balanced chemical equation</th>
<th>(2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number of particles (molecules)</td>
<td>(2 \text{ molecules} \text{H}_2(g) + 1 \text{ molecule} \text{O}_2(g) \rightarrow 2 \text{ molecules} \text{H}_2\text{O}(l))</td>
</tr>
<tr>
<td>Amount (mol)</td>
<td>(2 \text{ mol} \text{H}_2(g) + 1 \text{ mol} \text{O}_2(g) \rightarrow 2 \text{ mol} \text{H}_2\text{O}(l))</td>
</tr>
<tr>
<td>Mass (g)</td>
<td>(4.04 \text{ g} \text{H}_2(g) + 32.00 \text{ g} \text{O}_2(g) \rightarrow 36.04 \text{ g} \text{H}_2\text{O}(l))</td>
</tr>
<tr>
<td>Total mass (g)</td>
<td>(36.04 \text{ g} \text{reactants} \rightarrow 36.04 \text{ g} \text{product})</td>
</tr>
</tbody>
</table>
Section 3.4 Summary

In this section, you learned how the mole connects the amount of a substance to its mass. The concept of the mole plays a major role in predicting how much of a given substance is consumed or produced in a chemical reaction. In later science or chemistry studies, you will learn more about the mole and its role in predicting quantities involved in chemical reactions.

Check Your Understanding

1. Write a paragraph to describe the mole and Avogadro's number, and explain their importance in chemistry.

2. One mole of atoms of any element contains $6.02 \times 10^{23}$ atoms. Explain why elements have different molar masses.

3. Find the molar mass of each of the following pure substances.
   (a) ozone, $O_3(g)$
   (b) carbon monoxide, $CO(g)$
   (c) methane, $CH_4(g)$
   (d) calcium carbonate, $CaCO_3(s)$
   (e) ammonium nitrate, $NH_4NO_3(s)$
   (f) aluminium sulfate, $Al_2(SO_4)_3(s)$

4. How many molecules are there in 1 mol of water?

5. Determine the mass of each of the following samples.
   (a) 3.52 mol silicon, $Si(s)$
   (b) 0.0035 mol barium, $Ba(s)$
   (c) 545.13 mol ammonium hydroxide, $NH_4OH(s)$
   (d) $8.45 \times 10^{-4}$ mol zinc phosphate, $Zn_3(PO_4)_2(s)$

6. Express the following as molar amounts.
   (a) 2.33 g hydrogen chloride, $HCl(g)$
   (b) 16.0 kg carbon dioxide, $CO_2(g)$
   (c) $1.03 \times 10^5$ g propane, $C_3H_8(g)$
   (d) 6.84 g sucrose, $C_{12}H_{22}O_{11}(s)$

7. Thinking Critically Chemical equations are sometimes written with fractions as coefficients. For example:
   $$C_2H_2(g) + \frac{5}{2}O_2(g) \rightarrow 2CO_2(g) + H_2O(g)$$
   (a) Demonstrate that the equation is balanced.
   (b) Does using fractions as coefficients make sense if you consider that the coefficients refer to the numbers of molecules involved in the reaction? Explain your answer.
   (c) Does using fractions make sense if you assume that the coefficients refer to moles of reactants and products? Explain your answer.
8. At the beginning of this unit, you learned that dyes can be obtained from natural sources such as plants. Dyes can also be produced synthetically. In 1856, chemistry student William Perkin accidentally discovered a way to produce a mauve dye synthetically. A dress that was dyed with the original mauve dye is shown in Figure 3.26.

(a) The mauve dye, known today as mauveine or aniline purple, contains various compounds. One of the compounds is \((C_{26}H_{23}N_{4})_2SO_4\). What is the molar mass of this compound?

(b) If you assume that mauveine consists exclusively of \((C_{26}H_{23}N_{4})_2SO_4\), what is the mass of \(5.8 \times 10^3\) mol?

(c) How many moles of mauveine are contained in 25 tonnes?

---

Figure 3.26 Perkin's mauve dye started a fashion craze. Because he made his discovery while he was trying to produce something entirely different, Perkin's invention is an example of the role of luck in scientific discovery. Luck, however, was not enough. Perkin needed to analyze what at first seemed to be a failed experiment.

William Henry Perkin was 18 years old when he produced the first synthetic dye — by accident. He was working with coal tar to try to obtain quinine, a substance used to treat malaria. His work resulted only in a dark sludge. When he processed the sludge, however, he found that it produced a purple substance that worked very well as a dye. He called the dye "mauve" and patented the process for making it. Clothing dyed with mauve quickly became extremely fashionable. In fact, the period after Perkin's invention is sometimes referred to as the "Mauve Decade."
Now that you have completed this chapter, try to do the following. If you cannot, go back to the sections indicated in parentheses after each part.

(a) Describe evidence that would suggest that a chemical change has taken place. (3.1)

(b) Describe how to distinguish between an exothermic reaction and an endothermic reaction. (3.1)

(c) State the law of conservation of energy. Describe how the law applies to both exothermic and endothermic chemical reactions. (3.1)

(d) Predict combinations of ions that will form an insoluble compound in water, using a solubility chart. (3.1)

(e) State the law of conservation of mass. How does the law apply to chemical reactions? (3.2)

(f) Explain why a balanced chemical equation is consistent with the law of conservation of mass. (3.2)

(g) Identify where you would place the term “thermal energy” in an equation for a reaction that is exothermic and for one that is endothermic. (3.3)

(h) Write the balanced chemical equations for specific examples of a formation reaction, a decomposition reaction, a single-replacement reaction, and a double-replacement reaction. (3.3)

(i) Define the term “hydrocarbon.” Give examples of at least two different hydrocarbons and their uses. (3.3)

(j) List the products of burning a hydrocarbon by complete combustion and by incomplete combustion. (3.3)

(k) Explain the significance of Avogadro’s number. (3.4)

(l) Explain how to calculate the molar mass of a compound, using an example. (3.4)

(m) Show how to determine the number of moles in a given mass of a substance. (3.4)

Prepare Your Own Summary

Summarize this chapter by doing one of the following. Use a graphic organizer (such as a concept map), produce a poster, or write a summary to include the key chapter concepts. Here are a few ideas to use as a guide:

- Identify the evidence for chemical reactions.
- Explain why chemists represent reactions using balanced chemical equations.
- Classify chemical reactions into five types.
- Explain why 1 mol of carbon and 1 mol of sodium contain the same number of particles but have different masses.
- Relate the law of conservation of mass to the mole concept.

- Explain how the law of conservation of energy relates to chemical reactions, such as the combustion reaction shown below.
Key Terms

chemical reaction
chemical change
reactant
product
precipitate
exothermic reaction
endothermic reaction

Understanding Key Concepts

Section numbers are provided if you need to review.

1. List five categories of evidence of chemical reactions, and give examples of each. (3.1)

2. Distinguish between an exothermic change and an endothermic change. Give one example of each. (3.1)

3. Explain the importance of a closed system in demonstrating the law of conservation of mass through experiment. (3.2)

4. Explain why a balanced chemical equation is consistent with the law of conservation of mass. (3.2)

5. In what circumstances can incomplete combustion occur, and why can it be dangerous? (3.3)

6. Distinguish between a formation reaction and a decomposition reaction. Give one example of each. (3.3)

7. Which of the following compounds are hydrocarbons? Briefly explain your answers. (3.3)
   (a) C₃H₈(g)
   (b) C₆H₁₂O₆(s)
   (c) H₂CO₃(s)
   (d) CH₃OH(ℓ)
   (e) C₆H₆(ℓ)

8. Describe the relationship between the mole and Avogadro's number. (3.4)

Developing Skills

9. Predict whether each of the following ionic substances will be soluble in water.
   (a) NaOH(s)
   (b) CuCl₂(s)
   (c) AgClO₃(s)
   (d) CH₃COO(s)
   (e) Pb(OH)₂(s)
   (f) SrSO₄(s)

10. Write the balanced chemical equation for the displacement of aluminium in aluminium chloride by magnesium.

11. Gold is usually found as small, uncombined pieces of metal. One way to extract gold is to dissolve it in a cyanide solution (a hazardous procedure!). The gold can be precipitated from the cyanide solution by adding a metal such as zinc:
    2Au(CN)₂(aq) + Zn(s) → 2Au(s) + Zn(CN)₄(aq)
    What type of reaction is this?

12. Write a balanced chemical equation for each of the following word equations. Use the description of energy changes to write the term “thermal energy” on the correct side of the equation.
    (a) Solid sulfur combines with oxygen gas to form sulfur dioxide gas. The reaction is exothermic.
    (b) Solid sulfur combines with oxygen gas exothermically to form sulfur trioxide gas.
17. As part of a laboratory procedure, a student must dissolve 0.10 mol of silver nitrate, AgNO₃(s), in 100 mL of distilled water. What mass of silver nitrate should the student add to the distilled water?

18. A student finds the mass of a small beaker to be 12.5 g. The student adds some sodium hydrogen carbonate, NaHCO₃(s), to the beaker and finds the mass of the NaHCO₃(s) and the beaker to be 18.1 g. What amount of NaHCO₃(s), in moles, did the student add to the beaker?

19. You begin with 15.0 g of sodium chloride. You use electrical energy to cause the sodium chloride to decompose into its constituent elements.

   (a) How many moles of sodium chloride did you begin with?
   (b) What do you predict will be the combined mass of the products? Explain your answer.

Critical Thinking

20. Suppose that a winter storm brings down electrical lines over a wide area. You are told that there will be no electricity for three or four days. It is bitterly cold outside and, unfortunately, your furnace cannot operate without electricity. Someone suggests placing towels around doors and windows to keep out drafts, and bringing in the barbecue as a source of heat. Why would this last action be very dangerous?
Over the past 50 years, countless industrial activities have taken place all over Earth. Unfortunately some of these activities have resulted in the contamination of soil. Researchers like Professor Selma Guigard and her students at the University of Alberta are working to perfect a process that will remove unwanted or even dangerous metals from soil. Thanks to their work, contaminated soil can be made safe and usable once more.

**Q** I’ve read about environmental companies using micro-organisms to clean soil. Does your research involve anything like that?

**A** No. Although it’s popular for some clean-up jobs, bioremediation is not terribly successful in removing metals. Often the metals are toxic to the organisms that are meant to clean them up. The process we are investigating is a solvent extraction technology. It is a process that has been tried by others in the past, but the result was not very satisfactory.

**Q** Why not?

**A** The solvent that was used in the process was usually an organic compound, such as methanol. It successfully cleaned the soil, but it left behind a residue that often made the soil useless for anything but landfill. As well, after the cleaning process, the contaminated solvent had to be disposed of safely.

**Q** Is the method you are researching able to eliminate those problems?

**A** Yes. We are using a different solvent to clean the soil — carbon dioxide, CO₂. It is more environmentally friendly than organic solvents, so there is no problem with disposal or unsafe residues. We’re not using CO₂ in its natural state, though. We’re using it in its supercritical fluid form. Supercritical CO₂ is the same thing some manufacturers use to remove caffeine from coffee beans.

**Q** What is supercritical CO₂?

**A** When we bring CO₂ up to a certain pressure and temperature, it becomes what is called a supercritical fluid. It behaves a little like a liquid and a little like a gas, so it is able to do the job we need it to do. Other researchers have already used CO₂ in its supercritical state to extract unwanted organic compounds from soil. The idea of using it to remove metals was first proposed around 1991. Our lab is one of only a handful in Canada that are pursuing the idea.

**Q** What does the cleaning process entail?

**A** A batch of contaminated soil is closed inside a reactor, along with something called a “chelating agent.” CO₂ is pumped in, and the pressure and temperature within the reactor are set. The supercritical fluid remains in contact with the soil for several minutes or hours.
The length of time depends on the amount of soil and the metal that is being removed. When the CO₂ is removed, depressurized, and brought to normal temperature, it changes from its supercritical fluid state back to a gas. Then the metal deposits just fall out of it. We’re left with CO₂, which is recycled. We’re also left with a small amount of highly concentrated metal, which we hope we can recycle some day.

Q: How exactly does the CO₂ get the metal to leave the soil?
A: The metal is usually in the soil in the form of a charged ion. Carbon dioxide is not charged. So the metal won’t dissolve in CO₂ on its own. That’s why we add the chelating agent. If the right chelate is added, the metal bonds with the chelate more than with whatever it is bonded to in the soil. Once the metal combines with the chelate, the compound has no charge. The compound dissolves in the CO₂ and flows out with it.

Q: What sort of metals can you extract using this process?
A: We’ve been using copper as our test metal. Once the process has been perfected, we hope to use it to remove such things as lead, mercury, cadmium, and chromium. One day, far down the road, it may even remove radioactive elements.

Q: What stage is your research at right now?
A: My students and I are beginning to work on designing a way to pump new batches of soil into the reactor without having to depressurize in between. The depressurizing and repressurizing for each load of soil is one of the most expensive parts of the process right now. For the moment, however, our research is focussed on determining the best chelating agent to use. We hope to do some trials with soil samples in the near future.

**EXPLORING Further**

**Personal Poisons: Removing Metals from the Blood**

Many of the same toxic, heavy metals that contaminate soil, water, and air are often found in the bloodstream of people who have been exposed to these contaminants. Lead and mercury are especially worrisome, because they directly and indirectly cause a host of serious (and sometimes deadly) health problems. Anyone can be exposed to lead. Common sources include old paint, soil, and even some toys. Sources of mercury include old paint, some kinds of fish, and certain vaccines and drugs prepared prior to the late 1990s. Chelation therapy is a proven course of action for removing heavy metals from blood. What is this therapy, and what does it involve? Why are there controversies surrounding chelation therapy, what are the viewpoints, and what is being done to resolve them? As a starting point, go to [www.mcgrawhill.ca/links/sciencefocus10](http://www.mcgrawhill.ca/links/sciencefocus10). Is there lead or other heavy metals contaminating this site? Why does it matter? How could you find out?
Some suggested areas of investigation:

- Does every antacid tablet of the same brand contain the same quantity of active ingredient?
- How do antacid tablets of different brands compare in terms of neutralizing ability? What does this have to do with the mass of the tablets or the nature of the active ingredient?
- Do antacid tablets of different brands contain the same active ingredient? If not, how could you compare the effectiveness of different brands?
- What are the products of the reaction between an antacid and an acid? (Can you isolate and test any or all of these products?)

Possible methods of investigation:

- Is a gas produced by the reaction you are working with? If so, how can you use what you know about the law of conservation of mass? What can you determine by comparing the mass of the reaction vessel and contents before and after the reaction?
- Does a neutralization reaction take place? If so, how can you use indicators and drop by drop titration to compare the effectiveness of active ingredients in antacids?

Safety Precautions

- Do not mix chemicals without your teacher’s knowledge and approval.
- You will be working with dilute acids and/or bases in this investigation. Remember that bases and acids are corrosive. If any acid or base contacts your skin, rinse the area with plenty of cold water and inform your teacher. Clean up any spills as your teacher directs.
- List additional appropriate safety precautions as you design your experiment.

Skill Focus

For tips on how to set up a controlled experiment, turn to Skill Focus 6.
Materials
Brainstorm a list of the apparatus and materials that will be most appropriate to answer your question. You may also need electronic resources, construction materials, and art materials for your final report.

Initiate and Plan
1. With your group, decide on an experimental question to investigate. You might need to do further research to decide on the question.

2. If it is appropriate to your question, formulate a hypothesis that provides an explanation that relates to your question. Write a testable prediction that states what you think you will observe in your particular experiment.

3. Brainstorm ways to test your prediction. Any of the lab skills you have learned in this unit and in earlier courses may be useful. For example, you may have experimented with acid-base neutralization in Grade 9.

4. Design an experiment to test your prediction. Use words and diagrams to explain your design. Refer to the Experimental Design Checklist on the right for help.

Perform and Record
5. Set up and perform your experiment. Carry out two or more trials. Make modifications to your experiment as necessary.

6. Gather and record data and observations as you conduct your experiment. Decide how to record and present your data in a clear format.

Analyze and Interpret
7. Draw conclusions based on the results of your experiment. Discuss your conclusions with your group.

8. Did your findings agree with your prediction? Are you able to use your hypothesis to explain your results? Explain.

9. Write up your findings in a laboratory report. Be sure to include the following:
   - Introduction
   - Hypothesis and/or prediction
   - Procedure (step by step), including a diagram
   - Data/Observations in the form of words combined with graphs, tables, etc.
   - Conclusions

10. Evaluate your experimental design.
    (a) How would you change your design if you were going to perform the experiment again?
    (b) What further questions arose as you carried out your experiment? How might you investigate these questions?

Experimental Design Checklist
1. Have you clearly stated the purpose of your experiment (the question you want to answer)?
2. Have you included any important background information you obtained from your initial research?
3. Have you written a testable prediction stating what you expect to happen, or a hypothesis tentatively explaining what will happen, or both?
4. Have you written a step-by-step procedure?
5. Did you make a complete list of all the materials you need?
6. Have you listed all appropriate safety precautions?
7. Have you identified the manipulated, responding, and controlled variables in your experiment?
8. Have you identified major assumptions and sources of error that you can think of in your design?
9. Did you repeat your experiment several times?
Now that you have completed Chapters 1, 2, and 3, you can assess how much you have learned by answering the following questions. Before you begin, you may find it useful to return to each Chapter at a Glance and each Chapter Review.

**True/False**

In your notebook, indicate whether each statement is true or false. Correct each false statement.

1. Ernest Rutherford provided evidence for the electron by bombarding gold foil with alpha particles.
2. An anion has more electrons than protons.
3. The atoms in diatomic molecules are held together by ionic bonds.
4. Metals usually react to form cations, not anions.
5. All binary ionic compounds containing the sodium ion are soluble in water.
6. Burning propane is an example of an endothermic reaction.
7. In a decomposition reaction, two or more substances react to produce one substance.
8. A sample of water containing $6.02 \times 10^{23}$ molecules has a mass of 36 g.
9. One mole of iron, Fe, has a mass of 55.85 g.
10. One mole of molecular oxygen, $O_2(g)$ contains $6.02 \times 10^{23}$ atoms of oxygen.

**Completion**

In your notebook, complete each statement with the correct term or phrase.

11. J.J. Thomson observed the behaviour of produced in gas discharge tubes to provide evidence for the existence of .
12. Bohr's theory of the atom says that electrons have certain allowed energies called .
13. A compound containing only two elements is called a .
14. In forming an ionic compound, a calcium atom tends to transfer to an atom of a element.
15. In a weak acid, a proportion of acid molecules in solution.
16. Neutralization is an example of a reaction.
17. The reaction of sodium with water is an example of a reaction.
18. When methane, $CH_4(g)$, burns in sufficient oxygen, the reaction is called a reaction.
19. Incomplete combustion of hydrocarbons results in the production of a toxic gas called .
20. One mole of carbon-12 atoms has a mass of exactly g.

**Matching**

In your notebook, copy the following descriptions. Beside each description, write the term from the bulleted list that best fits the description. A term may be used once, more than once, or not at all.

- (a) common name of a strong acid
- (b) reaction that absorbs energy from its surroundings
- (c) group of atoms with a net charge; the atoms are joined by covalent bonds
- (d) type of bond that forms between two non-metal atoms
- (e) type of bond in a molecule of oxygen, $O_2$
- (f) unit for the amount of a substance
- (g) intermolecular bonds in water
• covalent bond
• mole (mol)
• gram (g)
• hydrogen bond
• ionic bond
• vinegar
• endothermic
• polyatomic ion
• muriatic acid

Multiple Choice
In your notebook, write the letter of the best answer for each of the following questions.

22. Which of the following is not a typical property of an ionic compound?
(a) Ionic crystals with a well-defined shape form at room temperature.
(b) Aqueous solutions of ionic compounds are good electrical conductors.
(c) Ionic solids conduct electricity.
(d) Ionic solids have relatively high melting points.
(e) Ionic solids are hard and brittle.

23. What are the formulas of the carbonate ion, the ammonium ion, and the nitrate ion?
(a) CO$_3^{2-}$, NH$_3^+$, and NO$_3^-$
(b) CO$_3^{2-}$, NH$_3^+$, and NO$_3^{2-}$
(c) CO$_3^-$, NH$_4^+$, and NO$_3^-$
(d) CO$_3^{2-}$, NH$_4^+$, and NO$_3^-$
(e) CO$_2^{3-}$, NH$_4^+$, and NO$_3^{2-}$

24. When the following equation is balanced, what is the value of $x$?
$w$Al + $x$H$_2$SO$_4$ $\rightarrow$ $y$Al$_2$(SO$_4$)$_3$ + $z$H$_2$
(a) 1
(b) 2
(c) 3
(d) 4
(e) 5

25. Which of the following ionic compounds is highly soluble?
(a) Ag$_2$SO$_4$(s)
(b) PbCl$_2$(s)
(c) CuCl$_2$(s)
(d) Ca(OH)$_2$(s)
(e) Ba(OH)$_2$(s)

Short Answer
26. Give two examples of anions that have the same number of electrons as neon, Ne.

27. Draw electron dot diagrams to represent an atom of each of the following:
(a) silicon (c) strontium
(b) sulfur (d) argon

28. Draw electron dot diagrams for each of the following pairs of elements. Then show how atoms of each pair could form a compound, and predict the formula of the compound.
(a) Si and H (c) Ca and N
(b) Be and O (d) Al and P

29. Name the following compounds.
(a) Fe$_2$(SO$_4$)$_3$(s) (c) SO$_3$(g)
(b) Pb(NO$_3$)$_4$(s) (d) HI(g)

30. Write formulas for the following compounds.
(a) copper(I) nitrate
(b) mercury(II) bromide
(c) nickel(III) sulfide
(d) calcium hydrogencarbonate

31. Write the molar mass of each of the following elements.
(a) sodium, Na
(b) barium, Ba
(c) silicon, Si
(d) iodine, I$_2$

32. Predict whether each of the following ionic compounds are soluble or insoluble in water.
(a) Ba(ClO$_3$)$_2$(s)
(b) NaCH$_3$COO(s)
(c) SrSO$_4$(s)
(d) NH$_4$OH(s)

33. What is the range of pH values for an acid?

34. Draw a concept map to explain how the average atomic mass of an element and its molar mass are related.
35. Explain how the law of conservation of mass is reflected in the coefficients in balanced chemical equations. Use an equation to illustrate your explanation.

36. Why do many household cleaning products carry warning labels? Why should you wear gloves when using an oven cleaner?

37. Explain why water is a polar molecule.

38. Choose two unusual properties of water.
   (a) State what practical consequence each property has.
   (b) Explain each property based on the structure of the water molecule.

**Problem Solving/Applying**

39. Copy each of the following word equations into your notebook. Then write the balanced chemical equation, and classify the reaction.
   (a) Sodium reacts with water to produce aqueous sodium hydroxide and hydrogen gas.
   (b) Solid magnesium oxide reacts with carbon dioxide to form magnesium carbonate.
   (c) When heated, solid ammonium chloride forms ammonia and hydrogen chloride gas.
   (d) Aqueous copper(II) sulfate reacts with solid nickel to form aqueous nickel(II) sulfate and copper metal.
   (e) Solid sodium carbonate reacts with sulfuric acid to form an aqueous solution of sodium sulfate, water, and carbon dioxide gas.
   (f) Ammonia gas reacts with hydrogen chloride gas to form solid ammonium chloride.

40. Copy each of the following skeleton equations into your notebook. Then balance the equation and indicate the type of reaction.
   (a) NaClO₃(s) → NaCl(s) + O₂(g)
   (b) NaOH(aq) + (NH₄)₂SO₄(aq) → Na₂SO₄(aq) + NH₃(g) + H₂O(l)
   (c) H₂SO₄(aq) + Al₂O₃(s) → Al₂(SO₄)₃(aq) + H₂O(l)
   (d) Zn(s) + FeCl₂(aq) → ZnCl₂(aq) + Fe(s)
   (e) C₃H₄(g) + H₂(g) → C₃H₈(g)
   (f) CaCO₃(s) + HCl(aq) → CaCl₂(aq) + H₂O(l) + CO₂(g)

41. Determine the molar mass of each of the following compounds.
   (a) chlorine dioxide, ClO₂(g)
   (b) disulfur dichloride, S₂Cl₂(g)
   (c) sodium nitrate, NaNO₃(s)
   (d) phosphorus pentabromide, PBr₅(g)
   (e) urea, CH₄N₂O(s)
   (f) iron(II) phosphate, Fe₃(PO₄)₂(s)

42. How many moles of each substance are in each of the following samples?
   (a) 34.1 g Au(s)
   (b) 1.08 g Cr₂O₃(s)
   (c) 4.33 × 10⁻² g NH₄Br(s)
   (d) 3.32 kg (NH₄)₂Cr₂O₇(s)

43. What is the mass of each of the following?
   (a) 3.70 mol H₂O(l)
   (b) 14.8 mol BaCrO₄(s)
   (c) 2 × 10³ mol HCl(g)
   (d) 0.0345 mol Fe₂O₃(s)

44. How many molecules of CO₂(g) are in 2 mol of CO₂(g)?

45. Suppose you had an oven that could reach extremely high temperatures. If you place a sample of malachite in the oven and heat it vigorously, would you expect to obtain copper? (The oven is made of an unreactive material.) Explain your answer.
46. Alloys of titanium with iron are used in airplanes and racing bicycles because of their strength and relatively low density. Titanium can be obtained by reacting TiCl4 at a high temperature with magnesium. The skeleton equation is:

\[ \text{TiCl}_4(\ell) + \text{Mg}(\ell) \rightarrow \text{Ti}(s) + \text{MgCl}_2(\ell) \]

(a) What is an alloy?
(b) Balance the equation and classify the reaction.
(c) At the temperature of the reaction, which metal has the higher melting point? Explain.

47. A student carries out a chemical reaction in the laboratory that generates a gas. The student decides to test the gas to identify it. The student lowers a burning splint into the gas, and the flame is immediately extinguished. The student concludes that the reaction produced the gas carbon dioxide. Do you agree with this conclusion? If so, explain why. If not, what other tests could you carry out to positively identify the gas?

48. The RNI (recommended nutrient intake) of iron for women is 0.0148 g per day. Ferrous gluconate, \( \text{Fe(C}_6\text{H}_{11}\text{O}_7)_{2(\ell)} \), is often used in iron supplements because it is easier than elemental iron for the body to absorb. Some iron-fortified breakfast cereals contain elemental iron metal as their source of iron.

(a) Determine the number of moles of elemental iron, \( \text{Fe} \), required by a woman each day according to the RNI.
(b) Determine the molar mass of ferrous gluconate.
(c) What mass of ferrous gluconate would satisfy the RNI for iron?
(d) Some evidence suggests that the elemental iron in iron-fortified breakfast cereals is absorbed only to a small extent. How could a breakfast cereal manufacturer assure that a consumer absorbs more iron? Make two suggestions, and list the pros and cons of each suggestion.
(e) Suggest a way you could test if a breakfast cereal contained elemental iron without looking at the list of ingredients.

49. Hemoglobin is a protein molecule in blood that carries oxygen to all of the cells of the body. Carbon monoxide is produced by the incomplete combustion of fuels. It is a poison because it can bind to hemoglobin in the place of oxygen, with a bond that is almost 200 times stronger than the hemoglobin-oxygen bond. If a person is working in the presence of carbon monoxide, poisoning may result.

(a) Based on the information above, how might a person with carbon monoxide poisoning be treated?
(b) Suggest at least two situations that would create a risk for carbon monoxide poisoning, and suggest ways to avoid or remedy those situations.

50. Mercury is highly toxic in its elemental form and in chemical compounds. Water-soluble mercury compounds such as mercury(II) nitrate, \( \text{Hg(NO}_3\text{)}_{2(\ell)} \), are particularly dangerous, because they are easily spread through an ecosystem via wastewater. One way to remove mercury(II) nitrate from wastewater is to add sodium sulfate, \( \text{Na}_2\text{S}(\ell) \).

(a) Predict what reaction will occur when mercury(II) nitrate and sodium sulfate react in aqueous solution. Write a balanced equation for the reaction and classify the reaction.
(b) Explain why this reaction would be useful as part of a process for removing mercury from wastewater.
(c) What would likely be the next step in the clean-up process?