Review Unit: Chemistry Review
“As a high school chemistry teacher, I have had opportunities to be creative, tell stories, play, learn, and teach chemistry in everyday life. I have explored the chemistry of pottery, food and cooking, and silver-smithing, and toured local industrial plants in the coal, aluminium, iron, and oil industries. I have enjoyed the taste and chemistry of wine, beer, and chocolate with fellow colleagues, and learned about the properties and characteristics that make each of these products so special.”

Peggy Au, Winston Churchill High School
Lethbridge, Alberta

Chemistry is everywhere, from the colourful Canada Day fireworks display to the ingredients in toothpaste and cosmetics at the local pharmacy and grocery store. “Chemical” is one of those words that people often associate with negative feelings or dangerous consequences. In fact, the comfortable lives we lead are due in large part to our understanding and application of chemistry. Some chemicals are harmful to people or the environment, but many are integral to life, such as the carbon dioxide, oxygen, water, and glucose in the cycle of photosynthesis and cellular respiration. The air we breathe and the food we eat are all based on the elements of the periodic table and their combinations. The water we drink must be processed chemically and physically before it is considered safe. Rocks and minerals form the foundation of the non-living environment. For example, magnesium carbonate from the Rocky Mountains becomes a main component of sidewalks in our bustling cities. Mined metal ores become pots, pans, building components, and jewellery. Copper pans cannot be used for cooking acidic foods such as tomatoes and lemon juice because they cause the copper to oxidize. Cookies baked without baking soda are as hard as rocks. Chemistry is discovery, and a sense of wonder about how and why elements interact and combine in the natural and human-made world. Chemistry surrounds us.
GENERAL OUTCOMES

In this unit, you will

• use atomic theory and the periodic table to classify, describe, explain, and predict the properties of the elements
• use atomic, ionic, and bonding theories to describe, explain, and predict the properties and chemical formulas for compounds
• use reaction generalizations to describe, explain, and predict simple chemical reactions
• describe the processes of science and the nature of scientific knowledge
• describe the differences between and interdependence of science and technology
• employ decision-making processes on science–technology–society issues
These questions will help you find out what you already know, and what you need to review, before you continue with this unit.

Knowledge

1. To facilitate the use of this textbook as a reference, answer the following questions. Use the periodic table and the Appendices to help you find the answers in the most efficient way. Include with your answer the textbook section where you found the answer.
   (a) What is the atomic number and melting point for the element copper?
   (b) From the Appendix, what is the chemical formula and recommended name for baking soda?
   (c) Sketch an Erlenmeyer flask.
   (d) Which variable should be listed on the vertical (y) axis of a graph?
   (e) Write the definition for “science” as found in the textbook.
   (f) What is the melting point of aluminium?
   (g) Describe the first step in the procedure for lighting a laboratory burner.
   (h) From the Appendix, list the headings for writing a laboratory report.
   (i) Describe the WHMIS symbol for a flammable chemical.

Skills

Refer to the Appendices to help you answer the following questions.

2. A student attempting to identify a pure substance from its density obtained the evidence shown in Table 1.
   (a) Construct and label a mass–volume graph from the evidence in Table 1.
   (b) From the graph, what mass of the substance has a volume of 12.7 mL?
   (c) From the graph, describe the relationship of mass and volume for this solid.

3. Imagine that a vacuum cleaner salesperson comes to your home to demonstrate a new model. The salesperson cleans a part of your carpet with your vacuum cleaner, and then cleans the same area again using the new model. A special attachment on the new model lets you see the additional dirt that the new model picked up. Analysis seems to indicate that the new model does a better job. Evaluate the experimental design and provide your reasoning.

4. List the manipulated and responding variables and one controlled variable of this experimental problem: “How does altitude affect the boiling point of pure water?”

5. Write an experimental design to answer the problem in question 4.

6. How must you dispose of the following substances in the laboratory?
   (a) broken beaker
   (b) corrosive solutions
   (c) toxic compounds
7. Draw a floor plan of the laboratory where you will be working. On your plan indicate the location of the following:
   (a) entrances (exits), including the fire exit
   (b) storage for aprons and eye protection
   (c) eyewash station
   (d) first-aid kit
   (e) fire extinguisher(s)
   (f) MSDS binders
   (g) container for broken glass

8. List the actions you should take if
   (a) your clothing catches fire
   (b) someone else’s clothing catches fire

9. Examine Figure 1. What safety rules are the students breaking?

10. (a) Identify the WHMIS symbols in Figure 2.
     (b) What should you do immediately if any chemical comes in contact with your skin?

11. Describe the procedure for lighting a burner by giving the correct sequence for the photographs in Figure 3. You may use a photograph more than once.
We could argue that chemistry is responsible for some of the hazards of modern life: environmental damage resulting from resource extraction; the toxic effects of some products; and the challenge of garbage disposal. However, that argument ignores the underlying truth: Chemistry has been fundamental to the development of society as we know it. We now have cleaner fuel, more durable and safer paints, easy-care clothing, inexpensive fertilizers, life-saving pharmaceuticals, corrosion-resistant tools and machinery, and unusual new materials that we are using in interesting new ways. Much of this innovation has made our lives better to some degree.

Chemistry is just another way to say “the understanding of the nature of matter.” Chemists through the ages and around the world have relied upon scientific inquiry, carrying out investigations and making careful observations. The periodic table sums up the results of many of those investigations and presents information about the elements (Figure 1). The observations that went into the creation of the periodic table also helped to create modern atomic theory. In turn, we can explain many of the patterns in the properties of the elements in terms of atomic theory. In this chapter, we will discuss the patterns used to classify elements and compounds, and consider how these patterns are explained by atomic theory.

**STARTING Points**

**Answer these questions as best you can with your current knowledge. Then, using the concepts and skills you have learned, you will revise your answers at the end of the chapter.**

1. Examine the periodic table on the inside front cover of this book. Identify and describe some similarities and differences between this table and the ones you used in previous grades.
2. Identify the parts of the atom and describe how they are arranged. According to this model, how do the atoms of the various elements differ from each other?
3. Identify and describe patterns in properties that you are aware of among the elements of the periodic table. Explain these patterns, using your model of the atom.
4. Why do elements form compounds? Use examples of compounds you are familiar with in your explanation.
5. Describe how the scientific community names and writes formulas for chemicals such as sodium, chlorine, table salt, sugar, and battery acid.

**Career Connections:**
Chemistry Teacher; Careers with Chemistry
Exploration

Combustion of Magnesium (Demonstration)

We know things in several different ways. We might see something happening with our own eyes, or we might take a measurement of some variable. These are qualitative and quantitative observations, respectively. Interpretations are statements that go beyond direct observation; for example, the magnesium reacted with oxygen.

In this demonstration you will watch a reaction and classify what you see as a qualitative or quantitative observation or an interpretation. The reaction is the burning of magnesium.

Only observe the burning of the magnesium when it is within the glass beaker. Never look directly at burning magnesium. The bright flame emits ultraviolet radiation that could harm your eyes. Due to possible reaction to bright light, persons known to have had seizures should not participate in the demonstration. Because of its hazardous nature, this demonstration should never be carried out by students.

Materials: lab apron, eye protection, rubber gloves, magnesium ribbon (5 cm), steel wool, laboratory burner and striker, crucible tongs, large glass beaker

- Observe the magnesium before, during, and after it is burned in air. Record all your observations.
- Take safety precautions, then light the laboratory burner (Appendix C).
- Use rubber gloves when handling the steel wool.
- Clean the magnesium ribbon with steel wool. Record any observations.
- Use tongs to hold the magnesium ribbon.
- Light the magnesium ribbon in the burner flame and hold the burning magnesium inside the glass beaker to observe.

(a) Classify your observations as qualitative or quantitative observations.
(b) Indicate which, if any, of your written statements are interpretations.

Figure 1
Since phosphorus (shown) spontaneously ignites in air, it must be stored in water. Other metals react with both oxygen and water and so must be stored in oil. Magnesium oxidizes more slowly in air, but can be ignited with a flame (see the Exploration below). Each element has its own set of physical and chemical properties.
1.1 Introduction: Science and Technology

What Is Science?
Science involves describing, predicting, and explaining nature and its changes in the simplest way possible. Scientists refine the descriptions of the natural world so that these descriptions are as precise and complete as possible. In science, reliable and accurate descriptions of phenomena become scientific laws.

In scientific problem solving, descriptions, predictions, and explanations are developed and tested through experimentation. In the normal progress of science, scientists ask questions, make predictions based on scientific concepts, and design and conduct experiments to obtain experimental answers. As shown in Figure 1, scientists evaluate this process by comparing the results they predicted with their experimental results.

Scientists make predictions that can be tested by performing experiments. Experiments that verify predictions lend support to the concepts on which the predictions are based. We try to explain events in order to understand them. Scientists, like young children, try to understand and explain the world by constructing concepts. Scientific explanations are refined to be as logical, consistent, and simple as possible.

Every investigation has a purpose—a reason why the experimental work is done. The purpose of scientific work is usually to create, test, and/or use a scientific concept. This order is chronological; for example, one scientist creates a concept, others test the concept, and, if the test is passed, many scientists (plus teachers and students) use the concept.

Scientific research is often very complex and involves, by analogy, many roadblocks, road repairs, and detours along the way. Scientists record all of their work, but the formal report submitted for publication often does not reflect the difficulties and circling back that is part of the process. The order of the report headings does not reflect a specific scientific method. Figure 1 illustrates the sections and sequence of a laboratory report. Depending on the purpose of the investigation, some sections (Hypothesis, Prediction, and Evaluation) may not be required. See Appendix B.

The Natures of Science and Technology
Science and technology are two different but parallel and intertwined human activities. **Science** is the study of the natural world with the goal of describing, explaining, and predicting substances and changes. The purpose of scientific investigations is to create, test, and/or use scientific concepts. **Technology** is the skills, processes, and equipment required to manufacture useful products or to perform useful tasks. Technology employs a systematic trial-and-error process whose goal is to get process or equipment to work. Technology generally does not seek scientific explanations for why it works.

Technology is an activity that runs parallel to science (Figure 2). Technology often leads science as in the development of processes for creating fire, cooking, farming, refining metal, and the invention of the battery. The use of fire, cooking, farming, and refining preceded the scientific understanding of these processes by thousands of years. The invention of the battery in 1800 was not understood scientifically until the early 1900s. Seldom does technology develop out of scientific research, although as the growth of scientific knowledge increases, the number of instances of technology as applied science is increasing.

The discovery of fire and the invention of the battery provided science with sources of energy with which to conduct experimental designs that would not otherwise have been
possible. Science would not progress very far without the increasingly advanced technologies available to scientists. Often scientific advances have to wait on the development of technologies for research to be done; for example, glassware, the battery, the laser, and the computer.

Often science is blamed for the effects of technology. Often people say, “Science did this,” or “Science did that.” Most often, though, it was a technological development that was responsible. Technologies and scientific concepts are created by people and used by people. We have to learn how to intelligently control and evaluate technological developments and scientific research, but we can’t unless we are scientifically and technologically literate.

Science is an intellectual pursuit founded on research. Concepts start as hypotheses (tentative explanations) and end as accepted or discarded theories or laws. Science meets its goal of concept creation by continually testing concepts. Old concepts are restricted, revised, or replaced when they do not pass the testing process. Scientific knowledge is one of many kinds of knowledge that is trusted and relied upon around the world. Scientists employ a skeptical attitude that results in scientific knowledge being constantly tested worldwide against the best evidence and logic available. Attitudes such as open-mindedness, a respect for evidence, and a tolerance of reasonable uncertainty are qualities found in a scientist. These attitudes represent a predisposition to act in a certain way, without claiming absolute knowledge.

What Is Chemistry?
Chemistry is the physical science that deals with the composition, properties, and changes in matter. Chemistry is everywhere around you, because you and your surroundings are composed of chemicals with a variety of properties. However, chemistry involves more than the study of chemicals. It also includes studying chemical reactions, chemical technologies, and their effects on the environment.

Chemistry is primarily the study of changes in matter. For example, coals burning, fireworks exploding, and iron rusting are all changes studied in chemistry. A chemical change or chemical reaction is a change in which one or more new substances with different properties are formed. Chemistry also includes the study of physical changes, such as water freezing to form ice crystals and boiling to form water vapour, during which no new substances are formed. (Physical changes are sometimes called phase changes or changes of state.)

Classifying Knowledge
Classification helps us to organize our knowledge. Classifying knowledge itself is an even more powerful tool. Here are some examples.

The Evidence section of an investigation report includes all observations related to a problem under investigation. An observation is a direct form of knowledge obtained by means of one of your five senses—seeing, smelling, tasting, hearing, or feeling. An observation might also be obtained with the aid of an instrument, such as a balance, a microscope, or a stopwatch.

Observations may also be classified as qualitative or quantitative. A qualitative observation describes qualities of matter or changes in matter; for example, a substance’s colour, odour, or physical state. A quantitative observation involves the quantity of matter or the degree of change in matter; for example, a measurement of the length or mass of magnesium ribbon. All quantitative observations include a number; qualitative observations do not.

An interpretation, which is included in the Analysis section of an investigation report, is an indirect form of knowledge that builds on a concept or an experience to further understand the observations.
describe or explain an observation. For example, observing the light and the heat from burning magnesium might suggest, based on your experience, that a chemical reaction is taking place. A chemist’s interpretation might be more detailed: The oxygen molecules collide with the magnesium atoms and remove electrons to form magnesium and oxide ions. Clearly, this statement is not an observation. The chemist did not observe the exchange of electrons.

Observable knowledge is called **empirical knowledge**. Observations are always empirical. **Theoretical knowledge**, on the other hand, explains and describes scientific observations in terms of ideas; theoretical knowledge is *not observable*. Interpretations may be either empirical or theoretical, and depend to a large extent on your previous experience of the subject. Table 1 gives examples of both kinds of knowledge.

<table>
<thead>
<tr>
<th>Table 1 Classification of Knowledge</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Type of knowledge</strong></td>
</tr>
<tr>
<td>empirical</td>
</tr>
<tr>
<td>theoretical</td>
</tr>
</tbody>
</table>

**Communicating Empirical Knowledge in Science**

Communication is an important aspect of science. Scientists use several means of communicating knowledge in their reports or presentations. Some ways of communicating empirical knowledge are presented below:

- **Empirical descriptions** communicate a single item of empirical knowledge, that is, an observation. For example, you might communicate the simple description that magnesium burns in air to form a white, powdery solid.
- **Tables of evidence** report a number of observations. The manipulated (independent) variable is usually listed first followed by the responding (dependent) variable. Table 2 shows results from a quantitative experiment that involved burning magnesium.

<table>
<thead>
<tr>
<th>Table 2 Mass of Magnesium Burned and Mass of Ash Produced</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Trial</strong></td>
</tr>
<tr>
<td>1</td>
</tr>
<tr>
<td>2</td>
</tr>
<tr>
<td>3</td>
</tr>
</tbody>
</table>

- **Graphs** are visual presentations of observations. According to convention, the manipulated variable is labelled on the x-axis, and the responding variable is labelled on the y-axis (Appendix F.4). For example, the evidence reported in Table 2 is shown as a graph in Figure 3. Graphs appear in the Analysis section of a lab report.
- **Empirical hypotheses** are preliminary generalizations that require further testing. Based on Figure 3, for example, you might tentatively suggest that the mass of the product of a reaction will always vary directly with the mass of a reacting substance.
- **Empirical definitions** are statements that define an object or a process in terms of observable properties. For example, a metal is a shiny, flexible solid.
• **Generalizations** are statements that summarize a limited number of empirical results. Generalizations are usually broader in scope than empirical definitions and often deal with a minor or sub-concept. For example, many metals slowly react with oxygen from the air in a process known as corrosion.

• **Scientific laws** are statements of major concepts based on a large body of empirical knowledge. Laws are more important and summarize more empirical knowledge than generalizations. For example, the burning of magnesium, when studied in greater detail (Table 3), illustrates the law of conservation of mass.

### Table 3  Mass of Magnesium, Oxygen, and Product of Reaction

<table>
<thead>
<tr>
<th>Trial</th>
<th>Mass of magnesium (g)</th>
<th>Mass of oxygen (g)</th>
<th>Mass of product (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>3.2</td>
<td>2.1</td>
<td>5.3</td>
</tr>
<tr>
<td>2</td>
<td>5.8</td>
<td>3.8</td>
<td>9.6</td>
</tr>
<tr>
<td>3</td>
<td>8.5</td>
<td>5.6</td>
<td>14.1</td>
</tr>
</tbody>
</table>

According to the evidence in Table 3, the total mass of magnesium and oxygen is generally equal to the mass of the product. Similar studies of many different reactions reflect the law of conservation of mass: In any physical or chemical change, the total initial mass of reactant(s) is equal to the total final mass of product(s).

For a statement to become accepted as a scientific law, evidence must first be collected from many examples and replicated by many scientists. Even after the scientific community recognizes a new law, that law is subjected to continuous experimental tests based on the ability of the law to describe, explain, and predict nature. Laws must accurately describe and explain current observations and predict future events in a simple manner.

### Section 1.1 Questions

1. Classify the following statements about carbon as observations or interpretations:
   (a) Carbon burns with a yellow flame.
   (b) Carbon atoms react with oxygen molecules to produce carbon dioxide molecules.
   (c) Global warming is caused by carbon dioxide.

2. Classify each of the following statements as one of the forms of empirical knowledge:
   (a) Elements are defined as substances that cannot be decomposed by heat or electricity.
   (b) The mass of products in a chemical reaction is always equal to the mass of reactants.
   (c) Graphite is a black powdery substance.
   (d) I believe that, in this case, the temperature will affect the rate of reaction.
   (e) Based upon the limited evidence available, the metals are all bendable.

3. Classify the following statements about carbon as empirical or theoretical:
   (a) Carbon in the form of graphite conducts electricity, whereas carbon in the form of diamond does not.
   (b) Graphite contains some loosely held electrons, whereas the electrons in diamond are all tightly bound in the atoms.

4. Scientific knowledge can be classified as empirical or theoretical.
   (a) What is the key distinction between these two types of knowledge?
   (b) How would you classify the knowledge in the Analysis section of an investigation report?

5. According to research by historians and philosophers, science and technology are related but different disciplines. Classify each of the following activities by Canadians as science or technology:
   (a) Aboriginal Canadians built birch bark canoes.
   (b) Harriet Brooks is the only person known to have worked in the laboratories of Ernest Rutherford, J.J. Thomson, and Marie Curie.
   (c) The telephone was invented by Alexander Graham Bell.
   (d) In 1990, Richard Taylor, who grew up in Medicine Hat, Alberta, was awarded the Nobel Prize for empirically testing and verifying the existence of quarks.

6. List the four characteristics of scientific communication.

7. Describe how a generalization differs from a scientific law.

8. Classify the following activities as science or non-science.
   (a) predicting the weather
   (b) fortune telling
   (c) astronomy
   (d) astrology
   (e) studying animal behaviour
   (f) observing the Northern Lights
   (g) ESP (extrasensory perception)
**Classifying Matter**

**Matter** is anything that has mass and occupies space. Anything that does not have mass or that does not occupy space—energy, happiness, and philosophy are examples—is not matter. To organize their knowledge of substances, scientists classify matter (Figure 1). A common classification differentiates matter as **pure substances**, whose composition is constant and uniform, and **mixtures**, whose composition is variable and may or may not be uniform throughout the sample. Empirically, **heterogeneous mixtures** are non-uniform and may consist of more than one phase. By analogy, your bedroom, for example, is a heterogeneous mixture because it consists of solids such as furniture, gases such as air, and perhaps liquids such as soft drinks. **Homogeneous mixtures** are uniform and consist of only one phase. Examples are alloys, tap water, aqueous solutions, and air.

**Figure 1**
An empirical classification of matter

<table>
<thead>
<tr>
<th>matter</th>
</tr>
</thead>
<tbody>
<tr>
<td>pure substances</td>
</tr>
<tr>
<td>constant composition</td>
</tr>
<tr>
<td>variable composition</td>
</tr>
<tr>
<td>mixtures</td>
</tr>
</tbody>
</table>

- **compounds** decomposed by heat and/or electricity
- **elements** not decomposed by heat or electricity
- **homogeneous mixtures** same properties throughout
- **heterogeneous mixtures** different properties in small samples

You can classify many substances as heterogeneous or homogeneous by making simple observations. However, some substances that appear homogeneous may, on closer inspection, prove to be heterogeneous (Figure 2). Introductory chemistry focuses on pure substances and homogeneous mixtures, commonly known as **solutions**.

This empirical (observable) classification system is based on the methods used to separate matter. The parts of both heterogeneous mixtures and solutions can be separated by physical means, such as filtration; distillation; chromatography; mechanically extracting one component from the mixture; allowing one component to settle; or using a magnet to separate certain metals. A pure substance cannot be separated by physical methods. A compound can be separated into more than one substance only by means of a chemical change involving heat or electricity. Separating a compound into its elements is called **chemical decomposition**. **Elements** cannot be broken down into simpler chemical substances by any physical or chemical means.

An **entity** is a general term that includes particles (sub-atomic entities such as protons, electrons, and neutrons), atoms, ions, molecules, and formula units. In this textbook, we restrict the use of the term “particle” to sub-atomic entities. Although the classification of matter is based on experimental work, theory lends support to this system. According to theory, elements are composed entirely of only one kind of atom. An **atom**, according to theory, is the smallest entity of an element that is still characteristic of that element. According to this same theory, **compounds** contain atoms...
of more than one element combined in a definite fixed proportion. Both elements and compounds may consist of molecules, distinct entities composed of two or more atoms. Solutions, unlike elements and compounds, contain entities of more than one substance, uniformly distributed throughout them.

A pure substance can be represented by a chemical formula, which consists of symbols representing the atoms present in the substance. You can use chemical formulas to distinguish between elements, which are represented by a single symbol, and compounds, which are represented by a formula containing two or more different symbols. Examples of formulas, along with empirical and theoretical definitions, are summarized in Table 1.

### Table 1 Definitions of Elements and Compounds

<table>
<thead>
<tr>
<th>Substance</th>
<th>Empirical definition</th>
<th>Theoretical definition</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>element</td>
<td>substance that cannot be broken down chemically into simpler units by heat or electricity</td>
<td>substance composed of only one kind of atom</td>
<td>Mg(s) (magnesium) O₂(g) (oxygen) C(s) (carbon)</td>
</tr>
<tr>
<td>compound</td>
<td>substance that can be decomposed chemically by heat or electricity</td>
<td>substance composed of two or more kinds of atoms</td>
<td>H₂O(l) (water) NaCl(s) (table salt) C₁₂H₂₂O₁₁(s) (sugar)</td>
</tr>
</tbody>
</table>

### Section 1.2 Questions

1. Describe how to distinguish experimentally between each of the following pairs of substances:
   (a) heterogeneous and homogeneous mixtures
   (b) solutions and pure substances
   (c) compounds and elements

2. The purpose of the investigation in this problem is to test the Design of decomposition by electricity to determine whether a pure substance is an element or a compound. Complete the Analysis and Evaluation of the investigation report. In your Evaluation, evaluate the Design only (Part 1; see Appendix B.2).

**Problem**

Are water and table salt classified as elements or compounds?

**Prediction**

According to current theoretical definitions of element and compound, as well as the given chemical formulas, water and table salt are classified as compounds. The chemical formulas indicate that water, H₂O(l), and table salt, NaCl(s), are composed of more than one kind of atom.

**Table 2 Passing Electricity through Samples**

<table>
<thead>
<tr>
<th>Sample</th>
<th>Description</th>
<th>Observations after passing electricity through sample</th>
</tr>
</thead>
<tbody>
<tr>
<td>water</td>
<td>colourless liquid</td>
<td>two colourless gases produced</td>
</tr>
<tr>
<td>molten table salt</td>
<td>colourless liquid</td>
<td>silvery solid and pale yellow-green gas formed (Figure 3)</td>
</tr>
</tbody>
</table>

3. Name two examples of each of the following:
   (a) pure substance
   (b) homogeneous mixture
   (c) heterogeneous mixture

4. Write a Design to determine whether a substance is an element or a compound. Make the Design extensive enough to provide a high degree of certainty in the answer.

5. John Dalton erroneously classified lime and several other compounds as elements because they would not decompose on heating. What does this indicate about the certainty of scientific knowledge?

**Figure 3**

(a) Sodium chloride (table salt)
(b) Sodium (dangerously reactive metal)
(c) Chlorine (poisonous, reactive gas)
1.3 Classifying Elements

Dmitri Mendeleev, a Russian chemist, created a periodic table in 1869. His periodic table communicated the periodic law—chemical and physical properties of elements repeat themselves in regular intervals, when the elements are arranged in order of increasing atomic number. (See the periodic table on the inside cover of this textbook.) A periodic table is a very useful tool upon which to base our chemical knowledge. The periodic table organizes the elements, for example, in groups and periods and as metals and nonmetals (Figure 1).

- A **family** or **group** of elements has similar chemical properties and includes the elements in a vertical column in the main part of the table.
- A **period** is a horizontal row of elements whose properties gradually change from metallic to nonmetallic from left to right along the row.
- Metals are located to the left of the “staircase line” in the periodic table, and nonmetals are located to the right. **Semi-metals**—sometimes called metalloids—are a class of elements that are distributed along the staircase line (Figure 2).

When chemists investigate the properties of materials, they must specify the conditions under which the investigations were carried out. For example, water is a liquid under normal conditions indoors, but in subzero winter temperatures, it becomes a solid outdoors. Ordinarily, tin is a silvery-white metal, but at temperatures below 13 °C, it gradually turns grey and crumbles easily. For the sake of accuracy and consistency, the International Union of Pure and Applied Chemistry (IUPAC), a governing body for scientific communication, has defined a set of standard conditions. Unless other conditions are specified, descriptions of materials are assumed to be at **standard ambient temperature and pressure**. Under these ambient (surrounding) conditions, known as SATP, the materials are at a temperature of 25 °C and a pressure of 100 kPa.

From many observations of the properties of elements, scientists have found that **metals** are shiny, bendable, and good conductors of heat and electricity. The majority

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**Figure 1**
The periodic table is divided into sections, each with commonly used names. The main group, or representative, elements are those in Groups 1, 2, and 12 to 18.

**Figure 2**
The properties of some elements have led to the creation of a class of elements called semi-metals.
of the known elements are metals, and all metals except mercury are solids at SATP. The remaining known elements are mostly nonmetals. Nonmetals are not shiny, not bendable, and generally not good conductors of heat and electricity in their solid form. At SATP, most nonmetals are gases and a few are solids. Solid nonmetals are brittle and lack the lustre of metals. Most nonmetals exist in compounds rather than in element form.

In the definitions for metals and nonmetals, bendable, ductile, and malleable are often used interchangeably, although they are different properties. Ductile means that the metal can be drawn (stretched) into a wire or a tube, and malleable means that the metal can be hammered into a thin sheet. The word lustrous is often used in these definitions in place of shiny.

**Groups of Elements**

Chemists classify elements, based on their similar physical and chemical properties, into (vertical) groups in our periodic table. Go to the link from the Nelson Web site and choose a vertical group to investigate. Report on similarities and differences among the elements of the group that you chose.

Periodic tables usually include each element’s symbol, atomic number, and atomic mass, along with other information that varies from table to table. The periodic table on the inside front cover features a box of data for each of the elements, and a key explaining the information in each box. The key is also shown in Figure 3. Note that theoretical data are listed in the column on the left, and empirically determined data are listed on the right.

**Figure 3**

This key, which also appears on this book’s inside front cover, helps you to determine the meaning of the numbers in the periodic table.

IUPAC specifies rules for chemical names and symbols. The IUPAC rules, which are summarized in many scientific references, are used all over the world. In Appendix I, you will find a list of all the English names of the elements in alphabetical order, along with their respective symbols.

**SUMMARY**

**IUPAC Rules for Element Symbols and Names**

- Element names should differ as little as possible among different languages. However, only the symbols are truly international.
- The first letter (only) of the symbol is always an uppercase letter (e.g., the symbol for cobalt is Co, not CO, co, or cO).
Families and Series of Elements

Some groups of elements have family and series names that are commonly used in scientific communication. It is important to learn these family and series names (Figure 1).

- The alkali metals are Group 1 elements. They are soft, silver-coloured metals that react violently with water to form basic solutions. The most reactive alkali metals are cesium and francium.
- The alkaline-earth metals are the Group 2 elements, not including beryllium and magnesium. They are light, reactive metals that form oxide coatings when exposed to air.
- The halogens are the elements in Group 17. They are all extremely reactive, with fluorine being the most reactive.
- The noble gases are the elements in Group 18. They are special because of their extremely low chemical reactivity.
- The main group elements are the elements in Groups 1, 2, and 12 to 18. Of all the elements, the main group elements best follow the periodic law.
- The transition elements are the elements in Groups 3 to 11. These elements exhibit a wide range of chemical and physical properties.

In addition to the common classes of elements described above, the bottom two rows in the periodic table also have common names. The lanthanoids are the elements with atomic numbers 58 to 71. The rare earth elements include the lanthanoids, and yttrium and scandium. The actinoids are the elements with atomic numbers 90 to 103. The synthetic (not naturally occurring) elements that have atomic numbers of 93 or greater are referred to as transuranic elements (beyond uranium).

### Section 1.3 Questions

1. What does the acronym IUPAC stand for?
2. Define SATP and state the reasons why IUPAC defined a set of standard conditions.
3. Classify the following chemicals as metals, semi-metals, or nonmetals:
   - (a) iron
   - (b) sulfur
   - (c) silicon
   - (d) gallium
4. List two physical and two chemical properties of the alkali metals.
5. Describe how the reactivity varies within the alkali metal family compared with the halogen family.
6. Nitrogen and hydrogen form a well-known compound, \( \text{NH}_3(g) \), ammonia. According to the position of phosphorus in the periodic table, predict the most likely chemical formula for a compound of phosphorus and hydrogen.
7. Complete the Prediction, Analysis, and Evaluation (Parts 2 and 3; see Appendix B.2) of the investigation report.

#### Purpose
The purpose of this investigation is to test the empirical definitions of metals and nonmetals.

#### Problem
What are the properties of the selected elements?

#### Design
Each element is observed at SATP, and the malleability and electrical conductivity are determined for the solid form.
Table 2 Properties of Selected Metals and Nonmetals

<table>
<thead>
<tr>
<th>Element</th>
<th>Appearance</th>
<th>Malleability of solid</th>
<th>Electrical conductivity of solid</th>
</tr>
</thead>
<tbody>
<tr>
<td>bromine</td>
<td>red-brown liquid</td>
<td>no</td>
<td>no</td>
</tr>
<tr>
<td>cadmium</td>
<td>shiny solid</td>
<td>yes</td>
<td>yes</td>
</tr>
<tr>
<td>chlorine</td>
<td>yellow-green gas</td>
<td>no</td>
<td>no</td>
</tr>
<tr>
<td>chromium</td>
<td>shiny solid</td>
<td>yes</td>
<td>yes</td>
</tr>
<tr>
<td>nickel</td>
<td>shiny solid</td>
<td>yes</td>
<td>yes</td>
</tr>
<tr>
<td>oxygen</td>
<td>colourless gas</td>
<td>no</td>
<td>no</td>
</tr>
<tr>
<td>platinum</td>
<td>shiny solid</td>
<td>yes</td>
<td>yes</td>
</tr>
<tr>
<td>phosphorus</td>
<td>white solid</td>
<td>no</td>
<td>no</td>
</tr>
</tbody>
</table>

8. Table 3 lists modern element symbols and ancient technological applications. Write the English IUPAC name for each element symbol.

Table 3 Ancient Technological Applications of Elements

<table>
<thead>
<tr>
<th>International symbol</th>
<th>Technological application</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sn</td>
<td>part of bronze (Cu and Sn) cutting tools, weapons, and mirrors used by Mayan and Inca civilizations</td>
</tr>
<tr>
<td>Cu</td>
<td>primary component of bronze and brass (Cu and Zn) alloys</td>
</tr>
<tr>
<td>Pb</td>
<td>used by Romans to make water pipes</td>
</tr>
<tr>
<td>Hg</td>
<td>a liquid metal used as a laxative by the Romans</td>
</tr>
<tr>
<td>Fe</td>
<td>produced by Egyptian iron smelters in 3000 BCE</td>
</tr>
<tr>
<td>S</td>
<td>burned for fumigation by Greeks in 1000 BCE</td>
</tr>
<tr>
<td>Ag</td>
<td>used in gold-silver alloys made by Greeks in 800 BCE</td>
</tr>
<tr>
<td>Sb</td>
<td>an element in ground ore used in early Egyptian cosmetics</td>
</tr>
<tr>
<td>Co</td>
<td>used in Egyptian blue-stained glass in 1500 BCE</td>
</tr>
<tr>
<td>Al</td>
<td>part of alum used as a fire retardant in 500 BCE</td>
</tr>
<tr>
<td>Zn</td>
<td>part of brass mentioned by Aristotle in 350 BCE</td>
</tr>
<tr>
<td>Fe</td>
<td>used by Inca and Olmec civilizations to make mirrors in 1000 BCE</td>
</tr>
</tbody>
</table>

9. Copy and complete Table 4. Note that the SATP states of matter are solid (s), liquid (l), and gas (g).

Table 4 Elements and Mineral Resources

<table>
<thead>
<tr>
<th>Mineral resource or use</th>
<th>Element name</th>
<th>Atomic number</th>
<th>Element symbol</th>
<th>Group number</th>
<th>Period number</th>
<th>SATP state</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a) high-quality ores at Great Bear Lake, NT</td>
<td>radium</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>(b) potash deposits in Saskatchewan</td>
<td></td>
<td>19</td>
<td>S</td>
<td>9</td>
<td>4</td>
<td></td>
</tr>
<tr>
<td>(c) extracted from Alberta sour natural gas</td>
<td>S</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>(d) radiation source for cancer treatment</td>
<td></td>
<td>9</td>
<td></td>
<td>4</td>
<td></td>
<td></td>
</tr>
<tr>
<td>(e) fuel in CANDU nuclear reactors</td>
<td></td>
<td>U</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>(f) mined in the Northwest Territories</td>
<td></td>
<td>14</td>
<td></td>
<td>2</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Extension

10. Search the Internet for the latest information on the discovery and naming of element 104 and beyond.
Theories and Atomic Theories

Empirical knowledge based on observation is the foundation for ideas in science. Usually, experimentation comes first and theoretical understanding follows. For example, the properties of some elements were known for thousands of years before a theoretical explanation was available.

So far in this chapter, you have encountered only empirical knowledge of elements, based on what has been observed. Curiosity leads scientists to try to explain nature in terms of what cannot be observed. This step—formulating ideas to explain observations—is the essence of theoretical knowledge in science. Albert Einstein (Figure 1) referred to theoretical knowledge as “free creations of the human mind.”

Communicating Theoretical Knowledge in Science

Scientists communicate theoretical knowledge in several ways:

- **Theoretical descriptions** are specific descriptive statements based on theories or models. For example, “a molecule of water is composed of two hydrogen atoms and one oxygen atom.”

- **Theoretical hypotheses** are ideas that are untested or extremely tentative. For example, “protons are composed of quarks that may themselves be composed of smaller particles.”

- **Theoretical definitions** are general statements that characterize the nature of a substance or a process in terms of a non-observable idea. For example, a solid is theoretically defined as “a closely packed arrangement of atoms, each atom vibrating about a fixed location in the substance.”

- **Theories** are comprehensive sets of ideas based on general principles that explain a large number of observations. For example, the idea that materials are composed of atoms is one of the principles of atomic theory; atomic theory explains many of the properties of materials. Theories are dynamic; they continually undergo refinement and change.

- **Analogies** are comparisons that communicate an idea in more familiar or recognizable terms. For example, an atom may be conceived as behaving like a billiard ball. All analogies “break down” at some level; that is, they have limited usefulness.

- **Models** are diagrams or apparatuses used to simplify the description of an abstract idea. For example, marbles in a vibrating box could be used to study and explain the three states of matter. Like analogies, models are always limited in their application.

Theories that are acceptable to the scientific community must describe observations in terms of non-observable ideas, explain observations by means of ideas, predict results in future experiments that have not yet been tried, and be as simple as possible in concept and application.

Dalton’s Atomic Theory

Recall that by the use of logic, the Greeks (Democritus) in about 300 BCE hypothesized that matter cut into smaller and smaller pieces would eventually reach what they called the atom—literally meaning indivisible. This idea was reintroduced over two thousand years later in 1805 by English chemist/schoolteacher John Dalton. Dalton re-created the modern theory of atoms to explain three important scientific laws—the laws of definite composition, multiple proportions, and conservation of mass.
Dalton’s model of the atom was that of a featureless sphere—by analogy, a billiard ball (Figure 2). Dalton’s atomic theory lasted for about a century, although it came under increasing criticism during the latter part of the 1800s.

**Creating Dalton’s Atomic Theory (1805)**

<table>
<thead>
<tr>
<th>Key experimental work</th>
<th>Theoretical explanation</th>
<th>Atomic theory</th>
</tr>
</thead>
<tbody>
<tr>
<td>Law of definite composition: Elements combine in a characteristic mass ratio.</td>
<td>Each atom has a particular combining capacity.</td>
<td>Matter is composed of indestructible, indivisible atoms, which are identical for one element, but different from other elements.</td>
</tr>
<tr>
<td>Law of multiple proportions: There may be more than one mass ratio.</td>
<td>Some atoms have more than one combining capacity.</td>
<td>Matter is composed of indestructible, indivisible atoms, which are identical for one element, but different from other elements.</td>
</tr>
<tr>
<td>Law of conservation of mass: Total mass remains constant (the same).</td>
<td>Atoms are neither created nor destroyed in a chemical reaction.</td>
<td>Matter is composed of indestructible, indivisible atoms, which are identical for one element, but different from other elements.</td>
</tr>
</tbody>
</table>

**Thomson’s Atomic Model**

Thomson’s model of the atom (1897) was a hypothesis that the atom was composed of electrons (negative particles) embedded in a positively charged sphere (Figure 3(a)). Thomson’s model of the atom is often communicated by using the analogy of a raisin bun (Figure 3(b)).

Table 2 summarizes the key experimental work that led to the creation of Thomson’s atomic theory, along with the theoretical explanations. Although you are not required to describe the experimental work, you do need to know that theories are created to explain evidence.

**Creating Thomson’s Atomic Theory (1897)**

<table>
<thead>
<tr>
<th>Key experimental work</th>
<th>Theoretical explanation</th>
<th>Atomic theory</th>
</tr>
</thead>
<tbody>
<tr>
<td>Arrhenius: the electrical nature of chemical solutions</td>
<td>Atoms may gain or lose electrons to form ions in solution.</td>
<td>Matter is composed of atoms that contain electrons (negatively charged particles) embedded in a positively charged material. The kind of element is characterized by the number of electrons in the atom.</td>
</tr>
<tr>
<td>Faraday: quantitative work with electricity and solutions</td>
<td>Particular atoms and ions gain or lose a specific number of electrons.</td>
<td>Matter is composed of atoms that contain electrons (negatively charged particles) embedded in a positively charged material. The kind of element is characterized by the number of electrons in the atom.</td>
</tr>
<tr>
<td>Crookes: qualitative studies of cathode rays</td>
<td>Electricity is composed of negatively charged particles.</td>
<td>Matter is composed of atoms that contain electrons (negatively charged particles) embedded in a positively charged material. The kind of element is characterized by the number of electrons in the atom.</td>
</tr>
<tr>
<td>Thomson: quantitative studies of cathode rays</td>
<td>Electrons are a component of all matter.</td>
<td>Matter is composed of atoms that contain electrons (negatively charged particles) embedded in a positively charged material. The kind of element is characterized by the number of electrons in the atom.</td>
</tr>
<tr>
<td>Millikan: charged oil drop experiment</td>
<td>Electrons have a specific fixed electric charge.</td>
<td>Matter is composed of atoms that contain electrons (negatively charged particles) embedded in a positively charged material. The kind of element is characterized by the number of electrons in the atom.</td>
</tr>
</tbody>
</table>
Rutherford's Atomic Theory

One of Thomson’s students, Ernest Rutherford (Figure 4), eventually showed that some parts of Thomson’s atomic theory were incorrect. Rutherford developed an expertise with nuclear radiation during the nine years he spent at McGill University in Montreal. Working with his team of graduate students, at Manchester in England, he devised an experiment to test Thomson’s model of the atom. The prediction, based on Thomson’s model, was that alpha particles should be deflected little, if at all. When some of the alpha particles were deflected at large angles and even backwards from the foil, the prediction was shown to be false, and Thomson’s model judged unacceptable (Figure 5). Rutherford created a nuclear model of the atom to explain the evidence gathered in this scattering experiment. The theoretical explanations for the evidence gathered are presented in Figure 6 and in Table 3.

**Prediction**

<table>
<thead>
<tr>
<th>alpha particles</th>
<th>metal foil</th>
</tr>
</thead>
</table>

**Evidence**

<table>
<thead>
<tr>
<th>alpha particles</th>
<th>metal foil</th>
</tr>
</thead>
</table>

Figure 4

Rutherford’s work with radioactive materials at McGill helped prepare him for his challenge to Thomson’s atomic theory.

Figure 5

Rutherford’s experimental observations were dramatically different from what he had expected based on Thomson’s model.

Figure 6

To explain his results, Rutherford suggested in his classic 1911 paper “that the atom consists of a central charge supposedly concentrated at a point.”

**DID YOU KNOW?**

Gold

The Incas used the malleability of gold to create many functional works of art. Gold was not viewed as an economic commodity, but was valued partly because it did not rust (oxidize). Due to its malleability, gold was used as the super-thin foil in Rutherford’s scattering experiment.

**WEB Activity**

Simulation—The Rutherford Scattering Experiment

Find out how Rutherford conducted his gold foil experiment.

www.science.nelson.com

**SUMMARY**

Creating Rutherford’s Atomic Theory (1911)

Table 3  Empirical Work Leading to Rutherford’s Atomic Theory

<table>
<thead>
<tr>
<th>Key experimental work</th>
<th>Theoretical explanation</th>
<th>Atomic theory</th>
</tr>
</thead>
<tbody>
<tr>
<td>Rutherford: A few positive alpha particles are deflected at large angles when fired at a gold foil.</td>
<td>The positive charge in the atom must be concentrated in a very small volume of the atom.</td>
<td>An atom is composed of a very tiny nucleus, which contains positive charges and most of the mass of the atom. Very small negative electrons occupy most of the volume of the atom.</td>
</tr>
<tr>
<td>Most materials are very stable and do not fly apart (break down).</td>
<td>A very strong nuclear force holds the positive charges within the nucleus.</td>
<td></td>
</tr>
<tr>
<td>Rutherford: Most alpha particles pass straight through gold foil.</td>
<td>Most of the atom is empty space.</td>
<td></td>
</tr>
</tbody>
</table>
Further research by several scientists led to creating the concepts of protons, neutrons, and isotopes. The evidence and explanations that expanded atomic theory are presented in Table 4.

### SUMMARY

#### Creating the Concepts of Protons, Neutrons, and Isotopes

<table>
<thead>
<tr>
<th>Key experimental work</th>
<th>Theoretical explanation</th>
<th>Atomic theory</th>
</tr>
</thead>
<tbody>
<tr>
<td>Soddy (1913): Radioactive decay suggests different nuclei of the same element.</td>
<td>Isotopes of an element have a fixed number of protons, but varying stability and mass (Figure 7).</td>
<td>Atoms are composed of protons, neutrons, and electrons. Atoms of the same element have the same number of protons and electrons, but may have a varying number of neutrons (isotopes of the element).</td>
</tr>
<tr>
<td>Rutherford (1914): The lowest charge on an ionized gas particle is from the hydrogen ion.</td>
<td>The smallest particle of positive charge is the proton.</td>
<td></td>
</tr>
<tr>
<td>Aston (1919): Mass spectrometer work indicates different masses for some atoms of the same element.</td>
<td>The nucleus contains neutral particles called neutrons.</td>
<td></td>
</tr>
<tr>
<td>Radiation is produced by bombarding elements with alpha particles.</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Still further empirical research with increasingly higher technologies allowed more precise determination of the relative masses of the subatomic particles. This more precise work confirmed that the electron and the proton had opposite charges, but that the charge was of the same magnitude (quantity) (Table 5).

The theoretical explanation for isotopes and the different masses of atoms of the same element is that atoms of elements can have a varying number of neutrons. Isotopes are designated by their mass number—the number of protons plus neutrons in their nucleus. The atomic number of an element could now be explained as the characteristic number of protons in the nucleus of atoms of that particular element. The number of neutrons can be calculated by subtracting the atomic number from the mass number. The atomic number and the mass number are shown in Figure 7.

### Bohr’s Atomic Theory

The genius of Niels Bohr lay in his ability to combine aspects of several theories and atomic models. He created a theory that, for the first time, could explain the periodic law. Bohr saw a relationship between the sudden end of a period in the periodic table and the quantum theory of energy proposed by German physicist Max Planck in 1900 and applied by Albert Einstein in 1905.

According to the Bohr atomic model, periods in the periodic table result from the filling of electron energy levels in the atom; for example, atoms in Period 3 have electrons in three energy levels. A period comes to an end when the maximum number of electrons is reached for the outer level. The maximum number of electrons in each energy level is given by the number of elements in each period of the periodic table; that is, 2, 8, 8, 18, etc. You may also recall that the last digit of the group number in the periodic table provides the number of electrons in the valence (outer) energy level. Although Bohr did his

---

**Canadian Diamonds**

Diamonds are the hardest mineral on Earth. Scientists indicate that diamonds from the Canadian north were formed 2.5 to 3.3 Ga (billion years) ago from carbon under high pressure and temperature at depths of 150 to 225 km. Volcanic action has brought diamonds closer to the surface in structures called kimberlite pipes. Once the pipes are discovered, they are drilled to test for the presence of diamonds. Buffalo Head Hills in Alberta is a diamond exploration area.

**DID YOU KNOW?**

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**Table 4** Experimental Work Leading to Theories of New Particles

<table>
<thead>
<tr>
<th>Key experimental work</th>
<th>Theoretical explanation</th>
<th>Atomic theory</th>
</tr>
</thead>
<tbody>
<tr>
<td>Soddy (1913): Radioactive decay suggests different nuclei of the same element.</td>
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<td></td>
</tr>
<tr>
<td>Radiation is produced by bombarding elements with alpha particles.</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Table 5** Masses and Charges of Subatomic Entities

<table>
<thead>
<tr>
<th>Particle</th>
<th>Relative mass</th>
<th>Relative charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>electron</td>
<td>1</td>
<td>1</td>
</tr>
<tr>
<td>proton</td>
<td>1836.12</td>
<td>1+</td>
</tr>
<tr>
<td>neutron</td>
<td>1838.65</td>
<td>0</td>
</tr>
</tbody>
</table>

**Figure 7**

Two isotopes of carbon. Carbon-12 is stable, but carbon-14 is radioactive. Carbon-14 is used for carbon dating of ancient artifacts.
Use the Bohr theory and the periodic table to draw energy-level diagrams for the phosphorus atom.

First, refer to the periodic table to find the position of phosphorus. Use your finger or eye to move through the periodic table from the top left along each period until you get to the element phosphorus. Starting with Period 1, your finger must pass through 2 elements, indicating that there is the maximum of 2 electrons in energy level 1. Moving on to Period 2, your finger moves through the full 8 elements, indicating 8 electrons in energy level 2. Finally, moving on to Period 3, your finger moves 5 positions to phosphorus, indicating 5 electrons in energy level 3 for this element.

The position of 2, 8, and 5 elements per period for phosphorus tells you that there are 2, 8, and 5 electrons per energy level for this atom. The information about phosphorus atoms in the periodic table can be interpreted as follows:

**Atomic number**: 15
- 15 protons and 15 electrons (for the atom)
**Period number**: 3
- Electrons in 3 energy levels
**Group number**: 15
- 5 valence electrons (the last digit of the group number)

To draw the energy-level diagram, work from the bottom up:
- **Sixth**, the 3rd energy level: 5 e⁻ (from group 15)
- **Fifth**, the 2nd energy level: 8 e⁻ (from eight elements in Period 2)
- **Fourth**, the 1st energy level: 2 e⁻ (from two elements in Period 1)
- **Third**, the protons: 15 p⁺ (from the atomic number)
- **Second**, the symbol: P (uppercase symbol from the table)
- **First**, the name of the atom: phosphorus atom (lowercase name)

Although the energy levels in this energy-level diagram are, for convenience, shown as equal distances apart, this is contrary to the evidence. Line spectra evidence indicates that higher energy levels are increasingly closer together (Figure 8).

### Practice

1. Draw an electron energy-level diagram for each of the following:
   (a) an atom of boron
   (b) an atom of aluminium
   (c) an atom of helium

### SAMPLE problem 1.1

Use the Bohr theory and the periodic table to draw energy-level diagrams for the phosphorus atom.

First, refer to the periodic table to find the position of phosphorus. Use your finger or eye to move through the periodic table from the top left along each period until you get to the element phosphorus. Starting with Period 1, your finger must pass through 2 elements, indicating that there is the maximum of 2 electrons in energy level 1. Moving on to Period 2, your finger moves through the full 8 elements, indicating 8 electrons in energy level 2. Finally, moving on to Period 3, your finger moves 5 positions to phosphorus, indicating 5 electrons in energy level 3 for this element.

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- **Third**, the protons: 15 p⁺ (from the atomic number)
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Although the energy levels in this energy-level diagram are, for convenience, shown as equal distances apart, this is contrary to the evidence. Line spectra evidence indicates that higher energy levels are increasingly closer together (Figure 8).
Formation of Monatomic Ions

In the laboratory, sodium metal and chlorine gas can react violently to produce a white solid, sodium chloride, commonly known as table salt (Figure 9). Sodium chloride is very stable and unreactive compared with the elements sodium and chlorine. Bohr originally suggested that the stable, unreactive behaviour of the noble gases was explained by their full outer electron orbits. According to this theory, when the neutral atoms collide, an electron is transferred from one atom to the other, and both atoms become entities called ions, which have an electrical charge (Figure 10). Sodium ions and chloride ions are monatomic ions—single atoms that have gained or lost electrons.

The theory of monatomic ion formation can be used to predict the formation of ions by most representative elements. However, the theory is restricted to these elements. Predictions cannot be made about:

- transition metals. Information about the ions of these elements can be obtained from the data in the periodic table on the inside front cover of this book.
- boron, carbon, and silicon. Experimental evidence indicates that these elements rarely form ions.
- hydrogen. Hydrogen atoms usually form positive ions by losing an electron. Although unusual, a negative hydrogen ion can be formed.
Positively charged ions are called **cations**. All of the monatomic cations are formed from the metallic and semi-metal elements when they lose electrons in an electron transfer reaction.

Negatively charged ions are called **anions**. All of the monatomic anions come from the nonmetallic and semi-metal elements. Names for monatomic anions use the stem of the English name of the element with the suffix “-ide” and the word “ion” (*Table 7*).

**Did You Know?**

**Useful Isotopes**

The following radioisotopes (radioactive isotopes) are produced artificially, for example, within the core of CANDU nuclear reactors. Most of the radioisotopes are used for medical diagnosis or therapy or for industrial or research work.

- Ir-192 — analysis of welds
- Co-60 — cancer treatment
- Te-99 — monitoring blood flow
- I-131 — hyperthyroid treatment
- C-14 — archeological dating
- Hg-203 — dialysis monitoring
- P-32 — white cell reduction
- Co-57 — monitoring vitamin B12
- Ti-201 — monitoring blood flow
- Sr-85 — bone scanning
- In-111 — brain tumor scanning
- Se-75 — pancreas tumor scanning

**Did You Know?**

**Analogy for Electron Transitions**

In an automobile, the transmission shifts the gears from lower to higher gears, such as from first to second, or downshifts from higher to lower gears. The gears are fixed: first, second, third. You cannot shift to “2½.” Similarly, electron energies in the Bohr model are fixed and electron transitions can only be up or down between specific energy levels.

**Evaluation of Scientific Theories**

“Physical concepts are free creations of the human mind, and are not, however it may seem, uniquely determined by the external world. In our endeavor to understand reality we are somewhat like a man trying to understand the mechanism of a closed watch. He sees the face and the moving hands, even hears its ticking, but he has no way of opening the case. If he is ingenious he may form some picture of a mechanism which could be responsible for all the things he observes, but he may never be quite sure his picture is the only one which could explain his observations. He will never be able to compare his picture with the real mechanism and he cannot even imagine the possibility of the meaning of such a comparison.”

It is never possible to prove theories in science. A theory is accepted if it logically describes, explains, and predicts observations. A major endeavour of science is to make predictions based on theories and then to test the predictions. Once the evidence is collected, a prediction may be

- **verified** if the evidence agrees within reasonable experimental error with the prediction. If this evidence can be replicated, the scientific theory used to make the prediction is judged to be acceptable, and the evidence adds further support and certainty to the theory;
- **falsified** if the evidence obviously contradicts the prediction. If this evidence can be replicated, the scientific theory used to make the prediction is judged to be unacceptable.

The ultimate authority in scientific work is the evidence (empirical knowledge) gathered during valid experimental work.

An unacceptable theory requires further action; there are three possible strategies.

- **Restrict** the theory. Treat the conflicting evidence as an exception and use the existing theory within a restricted range of situations.
- **Revise** the theory. This option is the most common. The new evidence becomes part of an improved theory.
- **Replace** the existing theory with a totally new concept.

These choices are often referred to as the three Rs.

### Table 8 Cation and Anion Formation

<table>
<thead>
<tr>
<th>Name</th>
<th>Atoms</th>
<th>Cations formed by metals</th>
<th>Anions formed by nonmetals</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nucleus</td>
<td>#p^+ = atomic number</td>
<td>#p^- = atomic number</td>
<td>#p^- = atomic number</td>
</tr>
<tr>
<td>Electrons</td>
<td>#e^- = #p^-</td>
<td>#e^- &lt; #p^-</td>
<td>#e^- &gt; #p^-</td>
</tr>
</tbody>
</table>

### Canadian Achievers—Harriet Brooks

Ontario-born Harriet Brooks (Figure 11) was Rutherford’s first graduate student at McGill University, Montreal. She was also the first woman at McGill to receive a Master of Science degree in physics. She made many contributions to science, including the discovery of radon as a radioactive by-product of radium. She was also the first person to realize that one element can change into another in a series of transformations. Research Brooks’s other research interests, and what life was like for a female scientist in the early 20th century.

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Figure 11
Harriet Brooks (1876–1933)
Section 1.4 Questions

1. What is the key difference between empirical and theoretical knowledge?

2. List four characteristics a theory must have to be accepted by the scientific community.

3. Describe the difference between a theory and a law.

4. According to the Bohr theory, what is the significance of full, outer electron orbits of atoms? Which chemical family has this unique property?

5. The alkali metals have similar physical and chemical properties (Figure 12). According to the Bohr theory, what theoretical similarity of alkali metal atoms helps to explain their empirical properties?

6. What is the ultimate authority in scientific work (what kind of knowledge is most trusted)?

7. Use the periodic table and theoretical rules to predict the number of occupied energy levels and the number of valence electrons for each of the following neutral atoms: beryllium, chlorine, krypton, iodine, lead, arsenic, and cesium.

8. If a scientific theory or other scientific knowledge is found to be unacceptable as a result of falsified predictions, what three options are used by scientists?

9. Write a theoretical definition of cation and anion.

10. The alkali metals all react violently with halogens to produce stable white solids. Draw energy-level diagrams (like Figure 10, page 24) for each of the following reactions:
   (a) lithium + chlorine → lithium chloride
   (b) potassium + fluorine → potassium fluoride

11. List the ion charges of the monatomic ions for the following families: alkali metals, alkaline-earth metals, Group 13, Group 15, Group 16, and halogens.

12. Draw energy-level diagrams for the following reactions between Group 2 elements and oxygen.
   (a) magnesium with oxygen (Figure 13)
   (b) calcium with oxygen

13. All the electron energy-level diagrams drawn in the previous questions have complete or filled outer energy levels. Describe the experimental evidence for these filled outer energy levels.

14. Write the symbols for the following atoms and ions (e.g., the sodium atom is Na, while the chloride ion is Cl⁻).
   (a) sulfur atom
   (b) oxide ion
   (c) lithium ion
   (d) phosphide ion
   (e) aluminium atom
   (f) gallium ion
   (g) rubidium ion
   (h) iodide ion

15. The purpose of this investigation is to test the theory of ions presented in this section. Complete the Prediction and Evaluation (Parts 2 and 3) of the investigation report.

   Problem
   What is the chemical formula of the compound formed by the reaction of aluminium and fluorine?

   Prediction
   According to the restricted quantum mechanics theory of atoms and ions, the chemical formula of the compound formed by the reaction of aluminium and fluorine is AlF₃.

   Design
   Aluminium and fluorine react in a closed vessel, and the chemical formula is calculated from the masses of reactants and products.

   Analysis
   According to the evidence gathered in the laboratory, the chemical formula of the compound formed by the reaction of aluminium and fluorine is AlF₃.

16. Describe some contributions Canadian scientists and/or scientists working in Canadian laboratories made to the advancement of knowledge about the nature of matter.

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Classifying Compounds

Before chemists could understand compounds, they had to devise ways to distinguish them from elements. Once they achieved this distinction, they could begin to organize their knowledge by classifying compounds. In this section, classification of compounds is approached in three different ways: by convention, empirically, and theoretically.

Classification of Compounds by Convention

Elements are commonly classified as metals or nonmetals. Given that compounds contain atoms of more than one kind of element, what combinations can result? Three classes of compounds are possible: metal–nonmetal, nonmetal–nonmetal, and metal–metal combinations (Figure 1).

- Metal–nonmetal combinations are called ionic compounds. An example is sodium chloride, NaCl (Figures 1 and 2).
- Nonmetal–nonmetal combinations are called molecular compounds. An example is sulfur dioxide, SO2.
- Metal–metal combinations are called alloys and inter-metallic compounds. Alloys include common metal–metal solutions (silver–gold alloys in coins). Inter-metallic compounds include CuZn, Cu5Zn8, and CuZn3 in brass at certain temperatures.

Empirical Classification of Ionic and Molecular Compounds

The properties of compounds can be used to classify compounds as ionic or molecular. Many properties are common to each of these classes, but by restricting the properties of each empirical definition to those easiest to identify, we can design diagnostic tests for

DID YOU KNOW?

To agree with the explanation of the empirically determined formula for sodium chloride, the model of a sodium chloride crystal must represent both a 1:1 ratio of ions and the shape of the salt crystal. The chemical formula, NaCl(s), represents one formula unit of the crystal, representing a 1:1 ratio of ions.

Figure 1
From two classes of elements, there can be three classes of compounds. This section of the chapter covers ionic and molecular compounds.

Figure 2
To agree with the explanation of the empirically determined formula for sodium chloride, the model of a sodium chloride crystal must represent both a 1:1 ratio of ions and the shape of the salt crystal. The chemical formula, NaCl(s), represents one formula unit of the crystal, representing a 1:1 ratio of ions.
Ionic and molecular compounds. A diagnostic test is a laboratory procedure conducted to identify or classify chemicals. Some of the common diagnostic tests used in chemistry are described in Appendix C.3.

**Empirical Definitions of Compounds**

In a series of replicated investigations scientists have found that ionic compounds are all solids at SATP. When dissolved in water, these compounds form solutions that conduct electricity. Scientists have also discovered that molecular compounds at SATP are solids, liquids, or gases that, when dissolved in water, form solutions that generally do not conduct electricity. These empirical definitions—a list of empirical properties that define a class of chemicals—will prove helpful throughout your study of chemistry. For example, electrical conductivity of a solution is an efficient and effective diagnostic test that determines whether a compound is ionic or molecular (Figure 3).

**Empirical Definition of Acids, Bases, and Neutral Compounds**

In science, it is not uncommon for new evidence to conflict with widely known and accepted theories, laws, and generalizations. Rather than viewing this as a problem, it is best regarded as an opportunity to improve our understanding of nature. As a result of the new evidence, the scientific concept is either restricted, revised, or replaced.

Not all compounds are either ionic or molecular. For example, aqueous hydrogen citrate (citric acid)—whose chemical formula is C₆H₈O₇(COOH)₃—is a compound composed of nonmetals. You might predict that this compound is molecular. However, a citric acid solution conducts electricity, which might lead you to predict that the compound is ionic (Figure 4). This conflicting evidence necessitates a revision of the classification system. A third class of compounds, called acids, has been identified, and the three classes together provide a more complete description of the chemical world.

**Acids** are solids, liquids, or gases as pure compounds at SATP that form conducting aqueous solutions that make blue litmus paper turn red. Acids exhibit their special properties only when dissolved in water. As pure substances, all acids, at this point in your chemistry education, have the properties of molecular compounds.

Experimental work has also shown that some substances make red litmus paper turn blue. This evidence has led to another class of substances: bases are empirically defined as compounds whose aqueous solutions make red litmus paper turn blue. Compounds whose aqueous solutions do not affect litmus paper are said to be neutral. These empirical definitions will be expanded in Chapter 6.

The properties of ionic compounds, molecular compounds, acids, and bases are summarized in Tables 1 and 2.

<table>
<thead>
<tr>
<th>Table 1</th>
<th>Properties of Ionic and Molecular Compounds and Acids</th>
</tr>
</thead>
<tbody>
<tr>
<td>State at SATP (pure substance)</td>
<td>Ionic</td>
</tr>
<tr>
<td>(s) only</td>
<td>(s), (l), or (g)</td>
</tr>
<tr>
<td>Conductivity of aqueous solution</td>
<td>high to low</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Table 2</th>
<th>Properties of Aqueous Solutions of Acids, Bases, and Neutral Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>Effect on blue litmus paper</td>
<td>Acidic</td>
</tr>
<tr>
<td>turns red</td>
<td>none</td>
</tr>
<tr>
<td>Effect on red litmus paper</td>
<td>none</td>
</tr>
</tbody>
</table>
Names and Formulas of Ionic Compounds

Communication systems in chemistry are governed by IUPAC. This organization establishes rules of communication to facilitate the international exchange of chemical knowledge. However, even when a system of communication is international, logical, precise, and simple, it may not be generally accepted if people prefer to use old names and are reluctant to change. There are many examples of chemicals that have both traditional names and IUPAC names. Chemical nomenclature is the systematic method for naming substances. Although names of chemicals are language-specific, the rules for each language are governed by IUPAC.

States of Matter in Chemical Formulas

Chemical formulas include information about the numbers and kinds of atoms or ions in a compound. It is also common practice in a formula to specify the state of matter: (s) to indicate “solid”; (l) to indicate “liquid”; (g) to indicate “gas”; and (aq) to indicate “aqueous,” which refers to solutions in water. Substances that readily form aqueous solutions are very soluble in water. Some examples of chemical formulas with states of matter are:

- NaCl(s) pure table salt
- CH₃OH(l) pure methanol
- O₂(g) pure oxygen
- C₁₂H₂₂O₁₁(aq) aqueous sugar solution

Predicting and Naming Ionic Compounds

Besides explaining empirically determined formulas, an acceptable theory must also be able to predict future empirical formulas correctly. Experimental evidence provides the test for a prediction made from a theory. A major purpose of scientific work is to test concepts by making predictions.

Binary Compounds

To predict an ionic formula from the name of a binary (two-element) compound, write the chemical symbol, with its charge, for each of the two ions. Then predict the simplest whole-number ratio of ions to obtain a net charge of zero. For example, for the compound aluminium chloride, the ions are Al³⁺ and Cl⁻. For a net charge of 0, the ratio of aluminium ions to chloride ions must be 1:3. The formula for aluminium chloride is, therefore, AlCl₃. This prediction agrees with the chemical formula determined empirically in the laboratory.

A complete chemical formula should also include the state of matter at SATP. Recall the generalization that all ionic compounds are solids at SATP. The complete formula is therefore, AlCl₃(s).

Al(s) + Cl₂(g) → Al³⁺Cl⁻₃(s) or AlCl₃(s)
aluminium + chlorine → aluminium chloride

The name of a binary ionic compound is the name of the cation followed by the name of the anion. The name of the metal ion is stated in full and the name of the nonmetal ion has an -ide suffix, for example, magnesium oxide, sodium fluoride, and aluminium sulfide. Remember, name the two ions.

Multi-Valent Metals

Most transition metals and some main group metals can form more than one kind of ion, that is, they are multi-valent. For example, iron can form an Fe³⁺ ion or an Fe²⁺ ion, although Fe³⁺ is more common. In the reaction between iron and oxygen, two
possible products form stable compounds. We can predict the chemical formulas for the possible ionic compounds formed by the reaction by examining ion charges and balancing charges—the total positive charge plus the total negative charge equals zero.

\[
\begin{align*}
2(3^{2-}) & \quad 3(2^{2+}) \\
Fe^{3+}O_2^{2-} & \quad Fe^{2+}O^{2-} \\
Fe_2O_3(s) & \quad FeO(s)
\end{align*}
\]

In the periodic table on this book’s inside front cover, selected ion charges are shown, with the most common (stable) charge listed first (Figure 5). If the ion of a multi-valent metal is not specified in a description or an exercise question, you can assume the charge on the ion is the most common one.

To name the compounds, name the two ions. In the IUPAC system, the name of the multi-valent metal includes the ion charge. The ion charge is given in Roman numerals in brackets; for example, iron(III) is the name of the Fe\(^{3+}\) ion and iron(II) is the name of the Fe\(^{2+}\) ion. The Roman numerals indicate the charge on the ion, not the number of ions in the formula. The names of the previously mentioned compounds are

\[
\begin{align*}
Fe(s) + O_2(g) & \rightarrow Fe_2O_3(s) \text{ iron(III) oxide} \\
Fe(s) + O_2(g) & \rightarrow FeO(s) \text{ iron(II) oxide}
\end{align*}
\]

### Compounds with Polyatomic Ions

Charges on polyatomic ions can be found in a table of polyatomic ions (see the inside back cover). Predicting the formula of ionic compounds involving polyatomic ions is done in the same way as for binary ionic compounds. Write the ion charges and then use a ratio of ions that yields a net charge of zero. For example, to predict the formula of a compound containing copper ions and nitrate ions, write the following:

\[
\begin{align*}
2^{2+} & \quad 2(1^{2-}) \\
Cu^{2+} & \quad Cu(NO_3)_2(s) \text{ copper(II) nitrate}
\end{align*}
\]

Two nitrate ions are required to balance the charge on one copper(II) ion (Figure 6). Note that parentheses are used in the formula to indicate the presence of more than one polyatomic ion. Do not use parentheses with one polyatomic ion or with simple ions. Do not write: Ag\(_3\)(SO\(_4\)_\(_2\))(s) or (Ag\(_2\))SO\(_4\)(s).

A **formula unit** of an ionic compound is a representation of the simplest whole number ratio of ions; for example, NaCl is a formula unit of sodium chloride. There is no such thing as a molecule of NaCl, only a formula unit. (See Figure 2, page 27.) The simplest ratio formula is also referred to as the **empirical formula**. All ionic formulas are empirical formulas.
Section 1.5

**Ionic Hydrates**

Empirical work indicates that some ionic compounds exist as hydrates; for example, white CuSO₄(s) also exists as blue CuSO₄•5H₂O(s) (Figure 7). You cannot predict the number of water molecules added to the ionic formula unit. You need to be given or to reference this information. The following examples illustrate the nomenclature of ionic hydrates that is recommended by IUPAC, with older nomenclature in parentheses.

- CuSO₄•5H₂O(s) is copper(II) sulfate—water (1/5) (copper(II) sulfate pentahydrate)
- Na₂CO₃•10H₂O(s) is sodium carbonate—water (1/10) (sodium carbonate decahydrate)
- Na₂SO₄•7H₂O(s) is sodium sulfate—water (1/7) (sodium sulfate heptahydrate)

**Figure 7**

Heating bluestone crystals, CuSO₄•5H₂O(s), produces a white powder, CuSO₄(s), according to the reaction

\[
\text{CuSO}_4 \cdot 5\text{H}_2\text{O}(s) + \text{heat} \rightarrow \text{CuSO}_4(s) + 5\text{H}_2\text{O}(g)
\]

Adding water to the white powder produces bluestone.

**DID YOU KNOW?**

Baking Soda

Baking soda (Figure 8) is one of the most versatile chemicals known. If you were to be deserted on an isolated island, baking soda would be a chemical of choice. Baking soda can be used for bathing, for brushing your teeth, for cleaning pots, for baking, and for extinguishing fires.

**Figure 6**

According to theory, two nitrate groups are required to balance the charge on one copper(II) ion. This theory agrees with observations.

**Learning Tip**

The convention for naming ionic hydrates has undergone a couple of changes in the last few decades. For example, copper(II) sulfate pentahydrate became copper(II) sulfate-5-water and now is copper(II) sulfate—water (1/5).

**SUMMARY Ionic Compounds**

Laboratory investigations indicate that there are classes of ionic compounds:
- binary ionic compounds such as NaCl, MgBr₂, and Al₂S₃
- polyatomic ionic compounds such as Li₂CO₃ and (NH₄)₂SO₄
- compounds of multi-valent metals such as CoCl₂ and CoCl₃

The empirically determined formulas of these types of compounds can be explained theoretically in a logically consistent way, using two concepts:
- Ionic compounds are composed of two kinds of ions: cations and anions.
- The sum of the charges on all the ions is zero.

Naming ionic compounds and writing ionic formulas:
- To name an ionic compound, name the two ions: first the cation and then the anion.
- To write an ionic formula, determine the ratio of ions that yields a net charge of zero.
Section 1.5 Questions

1. Distinguish, empirically and theoretically, between an element and a compound.

2. Distinguish, empirically and theoretically, between a metal and a nonmetal.

3. Classify each of the following as an element or compound:
   (a) C6H12O6(s) (glucose)   (c) CO(g) (poisonous)
   (b) Fe(s) (in steel)   (d) oxygen (20% of air)

4. Classify each of the following as a metal or nonmetal:
   (a) lead (poisonous)   (c) Hg(l)
   (b) chlorine (poisonous)   (d) Br2(l)

5. Classify each of the following as ionic or molecular:
   (a) C3H6O3(s) (glucose)   (b) Fe2O3 • 3H2O(s) (rust)
   (c) H2O(l) (water)   (d) potassium chloride (fertilizer)

6. Once empirical definitions of compounds are established, what kind of knowledge about compounds is likely to follow?

7. State the general names given to a positive ion and a negative ion.

8. Write the symbol, complete with charge, for each of the following ions:
   (a) chloride (b) chlorate (c) nitride (d) iron(III)
   (e) ammonium (f) hydroxide

9. Use IUPAC rules to name the following binary ionic compounds:
   (a) lime, CaO(s)   (b) road salt, CaCl2(s)   (d) a hydride, CaH2(s)
   (c) potash, KCl(s)

10. Write the chemical formulas and IUPAC names for the binary ionic products of the following chemical reactions. Do not get distracted by the formulas for the nonmetals or try to balance the equations. For example,
    Li(s) + Br2(l) → Li+Br− (s) or LiBr(s) (lithium bromide)
    (a) Sr(s) + O2(g) → (b) Ag(s) + Sn(s) →

11. Write the chemical formula and IUPAC name of the most common ionic product for each of the following chemical reactions. Do not get distracted by the formulas for the nonmetals or try to balance the equations. For example,
    Bi(s) + O2(g) → Bi2O3(s) (bismuth(III) oxide)
    (a) Ni(s) + O2(g) → (b) Pb(s) + Sn(s) →
    (c) Sn(s) + 1/2I2(s) → (d) Fe(s) + O2(g) →

12. Sketch diagrams of the sulfate and carbonate polyatomic ions.

13. Write empirical and theoretical definitions of an ionic compound.

14. For the IUPAC chemical names in each of the following word equations, write the corresponding chemical formulas (including the state at SATP) to form a chemical equation. (It is not necessary to balance the chemical equation.)
    (a) Sodium hypochlorite is a common disinfectant and bleaching agent. This compound is produced by the reaction of chlorine, Cl2(g), with lye:
       chlorine gas + aqueous sodium hydroxide → aqueous sodium chloride + liquid water + aqueous sodium hypochlorite
    (b) Sodium hypochlorite solutions are unstable when heated and slowly decompose:
       aqueous sodium hypochlorite → aqueous sodium chloride + aqueous sodium carbonate
    (c) The calcium oxide produced in the following reaction is used in a further reaction to produce oxalic acid, a common rust remover:
       aqueous sodium oxalate + solid calcium hydroxide → solid calcium oxalate + aqueous sodium hydroxide

15. Predict the international chemical formulas with states at SATP for the compounds formed from the following elements. Unless otherwise indicated, assume that the most stable metal ion is formed. (Write the full chemical equations, but do not balance the equations.) Also write the IUPAC name of the product.
    (a) Mg(s) + O2(g) → (b) Ba(s) + S4(s) → (c) Sc(s) + F2(g) →
    (d) Fe(s) + O2(g) → (e) Hg(l) + Cl2(g) → (f) Pb(s) + Br2(l) → (g) Co(s) + I2(s) →

16. For the chemical formulas in each of the following equations, write the corresponding IUPAC names to form a word equation. (Refer to the table of polyatomic ions.)
    (a) The main product of the following reaction (besides table salt) is used as a food preservative:
       NH4Cl(aq) + NaC6H5COO(aq) → NH4C6H5COO(aq) + NaCl(aq)
    (b) Aluminium compounds, such as the one produced in the following reaction, are important constituents of cement:
       Al(NO3)3(aq) + Na2SiO3(aq) → Al2SiO3(s) + NaNO3(aq)
    (c) Sulfides are foul-smelling compounds that can react with water to produce basic solutions:
       Na2S(s) + H2O(l) → NaHS(aq) + NaOH(aq)
    (d) Nickel(II) fluoride may be prepared by the reaction of nickel ore with hydrofluoric acid:
       NiO(s) + HF(aq) → NiF2(aq) + H2O(l)

17. Use current IUPAC rules to name the following hydrates. An older name is provided for reference, if possible.
    (a) FeSO4·7H2O(s) (ferrous sulfate heptahydrate)
    (b) Ni(NO2)2·6H2O(s) (nickelous nitrate hexahydrate)
    (c) Al2(SO4)3·18H2O(s) (aluminium sulfate-18-water)
    (d) 3CdSO4·8H2O(s) (no name possible under older systems of nomenclature for hydrates)

Extension

18. Why do systems of nomenclature change over time?

19. Why is it important to have internationally accepted systems of communication?
Many molecular formulas, such as H\textsubscript{2}O, NH\textsubscript{3}, and CH\textsubscript{4}, had been determined empirically in the laboratory by the early 1800s, but chemists could not explain or predict molecular formulas using the same theory as for ionic compounds. The theory that was accepted for these compounds was the idea that nonmetal atoms share electrons and that the sharing holds the atoms together in a group called a molecule. The chemical formula of a molecular substance—called a molecular formula—indicates the number of atoms of each kind in a molecule (Figure 1).

### Molecular Elements

As you have seen from the given chemical formulas for elements in the preceding examples and exercises, the chemical formula of all metals is shown as a single atom, whereas nonmetals frequently form diatomic molecules (i.e., molecules containing two atoms). Some useful rules are provided in Table 1. (Memorize the examples in this table.) An explanation of these rules is given in Unit 1. The diatomic elements end in -gen; for example, hydrogen, nitrogen, oxygen, and the halogens. O\textsubscript{3}(g) is a special unstable form of oxygen called ozone. S\textsubscript{8}(s) is called cyclooctasulfur, octasulfur, or usually just sulfur. Figure 2 (on the next page) illustrates models of some of these molecules.

#### Table 1 Chemical Formulas of Metallic and Molecular Elements

<table>
<thead>
<tr>
<th>Class of elements</th>
<th>Chemical structure</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>metallic elements</td>
<td>all are monatomic</td>
<td>Na(s), Hg(l), Zn(s), Pb(s)</td>
</tr>
<tr>
<td>molecular elements (nonmetallic)</td>
<td>some are diatomic</td>
<td>H\textsubscript{2}(g), N\textsubscript{2}(g), O\textsubscript{2}(g), F\textsubscript{2}(g), Cl\textsubscript{2}(g), Br\textsubscript{2}(l), I\textsubscript{2}(s)</td>
</tr>
<tr>
<td></td>
<td>some have molecules containing more than two atoms</td>
<td>O\textsubscript{3}(g), P\textsubscript{4}(s), S\textsubscript{8}(s)</td>
</tr>
<tr>
<td></td>
<td>all noble gases are monatomic</td>
<td>He(g), Ne(g), Ar(g)</td>
</tr>
<tr>
<td>other elements</td>
<td>the rest of the elements can be assumed to be monatomic</td>
<td>C(s), Si(s)</td>
</tr>
</tbody>
</table>

### Molecular Compounds

The names of some compounds communicate the number of atoms in a molecule. IUPAC has assigned Greek numerical prefixes to the names of molecular compounds formed from two different elements (Table 2). Other naming systems are used when a molecule has more than two kinds of atoms.

The following are examples of names of binary molecular compounds. Recall that binary refers to compounds composed of only two kinds of atoms and that molecular refers to compounds composed only of nonmetals.

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Product</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>C(s) + S\textsubscript{8}(s) → CS\textsubscript{2}(l)</td>
<td>carbon disulfide</td>
<td></td>
</tr>
<tr>
<td>N\textsubscript{2}(g) + I\textsubscript{2}(s) → NI\textsubscript{3}(s)</td>
<td>nitrogen triiodide</td>
<td></td>
</tr>
<tr>
<td>N\textsubscript{2}(g) + O\textsubscript{2}(g) → NO\textsubscript{2}(g)</td>
<td>nitrogen dioxide</td>
<td></td>
</tr>
<tr>
<td>P\textsubscript{4}(s) + O\textsubscript{2}(g) → P\textsubscript{4}O\textsubscript{10}(s)</td>
<td>tetraphosphorus decaoxide</td>
<td></td>
</tr>
</tbody>
</table>

You will predict the chemical formulas for molecular compounds in Unit 1.
Chapter 1

NEL

Naming Molecular Compounds

According to IUPAC rules, the prefix system is used only for naming binary molecular compounds—molecular compounds composed of two kinds of atoms.

For hydrogen compounds such as hydrogen sulfide, $\text{H}_2\text{S}(\text{g})$, the common practice is not to use the prefix system. In other words, we do not call this compound dihydrogen sulfide.

The molecular formulas and names of many molecular compounds must be memorized, referenced, or given. Some common molecular compounds whose names and formulas should be memorized are given in Table 3.

Figure 2

Models representing the molecular elements $\text{P}_4(\text{s})$, $\text{S}_8(\text{s})$, $\text{H}_2(\text{g})$, and $\text{Br}_2(\text{l})$.

Table 3

<table>
<thead>
<tr>
<th>IUPAC name</th>
<th>Molecular formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>water</td>
<td>$\text{H}_2\text{O}(\text{l})$ or $\text{HOH}(\text{l})$</td>
</tr>
<tr>
<td>hydrogen peroxide</td>
<td>$\text{H}_2\text{O}_2(\text{l})$</td>
</tr>
<tr>
<td>ammonia</td>
<td>$\text{NH}_3(\text{g})$</td>
</tr>
<tr>
<td>glucose</td>
<td>$\text{C}<em>6\text{H}</em>{12}\text{O}_6(\text{s})$</td>
</tr>
<tr>
<td>sucrose</td>
<td>$\text{C}<em>12\text{H}</em>{22}\text{O}_{11}(\text{s})$</td>
</tr>
<tr>
<td>methane</td>
<td>$\text{CH}_4(\text{g})$</td>
</tr>
<tr>
<td>propane</td>
<td>$\text{C}_3\text{H}_8(\text{g})$</td>
</tr>
<tr>
<td>octane</td>
<td>$\text{C}<em>8\text{H}</em>{18}(\text{l})$</td>
</tr>
<tr>
<td>methanol</td>
<td>$\text{CH}_3\text{OH}(\text{l})$</td>
</tr>
<tr>
<td>ethanol</td>
<td>$\text{C}_2\text{H}_5\text{OH}(\text{l})$</td>
</tr>
<tr>
<td>hydrogen sulfide</td>
<td>$\text{H}_2\text{S}(\text{g})$</td>
</tr>
</tbody>
</table>

SUMMARY

Elements and Molecular Compounds

- Empirically, molecular compounds as pure substances are solids, liquids, or gases at SATP. If they dissolve in water, their aqueous solutions do not conduct electricity.
- Theoretically, molecular elements and compounds are formed by nonmetal atoms bonding covalently to share electrons in an attempt to obtain the same number of electrons as the nearest noble gas.
- All metallic elements are monatomic; for example, aluminium is $\text{Al}(\text{s})$ and iron is $\text{Fe}(\text{s})$.
- The chemical formulas for nonmetallic elements should be memorized from Table 1 on page 33.
- The chemical formulas and/or the names of molecular compounds are given. You will predict these formulas in Chapters 8 and 9.
- Memorize the prefixes provided in Table 2 on page 33.
- Memorize the chemical formulas, names, and states of matter at SATP for common binary and ternary molecular compounds in Table 3. For other molecular compounds referred to in questions, you are given the states of matter.
- The chemical formulas for most binary molecular compounds are obtained from the prefixes in the given names; for example, dinitrogen tetraoxide gas is $\text{N}_2\text{O}_4(\text{g})$.
- The chemical names for most binary molecular compounds use prefixes to communicate the formula subscripts; for example, $\text{N}_2\text{S}_3(\text{l})$ is dinitrogen pentasulfide.
- The SATP states of matter of metallic and nonmetallic elements are memorized or referenced from the periodic table.

Naming and Writing Formulas for Acids and Bases

In this chapter, acids and bases are given very restricted empirical and theoretical definitions. Aqueous hydrogen compounds that make blue litmus paper turn red are classified as acids and are written with the hydrogen appearing first in the formula. For example, $\text{HCl}(\text{aq})$ and $\text{H}_2\text{SO}_4(\text{aq})$ are acids. $\text{CH}_4(\text{g})$ and $\text{NH}_3(\text{g})$ are not acids, so hydrogen is written last in the formula. In some cases, hydrogen is written last if it is part of a group such as the $\text{COOH}$ group; for example, $\text{CH}_3\text{COOH}(\text{aq})$. These $-\text{COOH}$ acids are organic acids and are described in more detail in Chapter 9.

Acids

Empirically, acids as pure substances are molecular compounds, as evident from their solid, liquid, and gas states of matter. Theoretically, they are composed of nonmetals that share
electrons. However, the formulas of acids can be explained and predicted by assuming that they are ionic compounds. For example, the chemical formulas for the acids HCl(aq), H₂SO₄(aq), and CH₃COOH(aq) can be explained as follows:

\[ \text{H}^{+}\text{Cl}^{-}(aq), \quad \text{H}^{+}\text{SO}_4^{2-}(aq), \quad \text{CH}_3\text{COO}^-\text{H}^+(aq) \]

The chemical formulas for acids can also be predicted by assuming that these aqueous molecular compounds of hydrogen are ionic:

aqueous hydrogen sulfide is \( \text{H}^{+}\text{S}^2-(aq) \), or \( \text{H}_2\text{S}(aq) \)
aqueous hydrogen sulfate is \( \text{H}^{+}\text{SO}_4^{2-}(aq) \), or \( \text{H}_2\text{SO}_4(aq) \)
aqueous hydrogen sulfite is \( \text{H}^{+}\text{SO}_3^{2-}(aq) \), or \( \text{H}_2\text{SO}_3(aq) \)

Acids are often named according to more than one system because they have been known for so long that the use of traditional names persists (Figure 3). IUPAC suggests that names of acids should be derived from the IUPAC name for the compound. In this system, sulfuric acid would be named aqueous hydrogen sulfate. However, the classical system of nomenclature is well entrenched, so it is necessary to know two or more names for many acids, especially the common ones.

The classical names for acids are based on anion names, according to three simple rules:

- If the anion name ends in “-ide,” the corresponding acid is named as a “hydro-—ic” acid. Examples are hydrochloric acid, HCl(aq), hydrosulfuric acid, H₂S(aq), and hydrocyanic acid, HCN(aq).
- If the anion name ends in “-ate,” the acid is named as a “—ic” acid. Examples are nitric acid, HNO₃(aq), sulfuric acid, H₂SO₄(aq), and phosphoric acid, H₃PO₄(aq).
- If the anion name ends in “-ite,” the acid is named as a “—ous” acid. Sulfurous acid, H₂SO₃(aq), nitrous acid, HNO₂(aq), and chlorous acid, HClO₂(aq), are examples.

The classical system of acid nomenclature is part of a system for naming a series of related compounds. Table 4 lists the acids formed from five different chlorine-based anions to illustrate this naming system.

<table>
<thead>
<tr>
<th>Formula</th>
<th>IUPAC name</th>
<th>Classical name</th>
<th>Systematic IUPAC name</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>ClO₄⁻</td>
<td>perchlorate ion</td>
<td>perchloric acid</td>
<td>aqueous hydrogen perchlorate</td>
<td>HClO₄(aq)</td>
</tr>
<tr>
<td>ClO₃⁻</td>
<td>chlorate ion</td>
<td>chloric acid</td>
<td>aqueous hydrogen chlorate</td>
<td>HClO₃(aq)</td>
</tr>
<tr>
<td>ClO₂⁻</td>
<td>chlorite ion</td>
<td>chloric acid</td>
<td>aqueous hydrogen chlorite</td>
<td>HClO₂(aq)</td>
</tr>
<tr>
<td>ClO⁻</td>
<td>hypochlorite ion</td>
<td>hypochlorous acid</td>
<td>aqueous hydrogen hypochlorite</td>
<td>HClO(aq)</td>
</tr>
<tr>
<td>Cl⁻</td>
<td>chloride ion</td>
<td>hydrochloric acid</td>
<td>aqueous hydrogen chloride</td>
<td>HCl(aq)</td>
</tr>
</tbody>
</table>

**Bases**

Chemists have discovered that all aqueous solutions of ionic hydroxides make red litmus paper turn blue; that is, these compounds are bases. Other solutions have been classified as bases, but for the time being, restrict your definition of bases to aqueous ionic hydroxides such as NaOH(aq) and Ba(OH)₂(aq). The name of the base is the name of the ionic hydroxide; for example, aqueous sodium hydroxide and aqueous barium hydroxide.
• Empirically, acids are aqueous molecular compounds of hydrogen that form electrically conductive solutions and turn blue litmus paper red.
• By convention, the formula for an empirically identified acid is written as $\text{H}\underline{\text{____}}(\text{aq})$ or $\text{COOH}(\text{aq})$.
• As pure substances, acids are molecular compounds, and, thus, can be solids, liquids, or gases; $\text{HCl}(\text{g})$, $\text{CH}_3\text{COOH}(\text{l})$, and $\text{C}_6\text{H}_5\text{OH}(\text{COOH})_3(\text{s})$.
• The chemical formulas and electrical conductivity of aqueous solutions of acids can be explained and predicted by assuming that these molecular compounds are ionic; for example, $\text{H}^+\text{SO}_4^{\text{-2}}(\text{aq})$ or $\text{H}_2\text{SO}_4(\text{aq})$.
• The classical names for acids follow this pattern: hydrogen ____ide becomes a “hydro______ ic” acid; hydrogen ______ate is a “_______ic” acid; hydrogen ______ite is a “_________ous” acid; and hydrogen hypo______ite is a “hypo________ous” acid.
• The IUPAC name for an acid is aqueous hydrogen ________; for example, aqueous hydrogen sulfate for $\text{H}_2\text{SO}_4(\text{aq})$.

Empirically, bases are aqueous ionic hydroxides that form electrically conductive solutions and turn red litmus paper blue.

There is no special nomenclature system for bases. They are named as ionic hydroxides; for example, aqueous hydrogen sulfate for $\text{H}_2\text{SO}_4(\text{aq})$.

• Empirically, bases are aqueous ionic hydroxides that form electrically conductive solutions and turn red litmus paper blue.

SUMMARY

Section 1.6 Questions

1. Until a theoretical way of knowing molecular formulas is available, you must be given the formula or name and then rely on memory or use the prefix system to provide the name or formula as required. Provide the names or formulas (complete with the SATP states of matter) for the following substances:
   (a) chlorine (toxic)
   (b) phosphorus (reacts with air)
   (c) $\text{C}_2\text{H}_5\text{OH}(\text{l})$ (alcohol)
   (d) methane (fuel)
   (e) helium (inert)
   (f) carbon (black)
   (g) $\text{NH}_3(\text{g})$ (smelling salts)

2. Write the chemical formulas for the following molecular substances emitted as gases from the exhaust system of an automobile. Some of these substances may produce acid rain:
   (a) carbon dioxide
   (b) carbon monoxide
   (c) nitrogen dioxide
   (d) sulfur dioxide
   (e) nitrogen
   (f) octane
   (g) nitrogen monoxide
   (h) dinitrogen oxide
   (i) dinitrogen tetroxide
   (j) water

3. Write unbalanced chemical equations to accompany the given statements or word equations, including the states at SATP. For example,
   
   nitrogen + oxygen $\rightarrow$ nitrogen dioxide
   $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g})$

   (a) Solid silicon reacts with gaseous fluorine to produce gaseous silicon tetrafluoride.

   (b) Solid boron reacts with gaseous hydrogen to produce gaseous diboron tetrahydride.

   (c) Aqueous sucrose and water react to produce aqueous diboron tetrahydride.

   (d) Methane gas reacts with oxygen gas to produce aqueous ethanol and carbon dioxide gas.

   (e) nitrogen + oxygen $\rightarrow$ nitrogen monoxide

   (f) nitrogen monoxide + oxygen $\rightarrow$ nitrogen dioxide

   (g) octane + oxygen $\rightarrow$ carbon dioxide + water vapour

   (h) octane + oxygen $\rightarrow$ carbon dioxide + carbon monoxide + carbon + water vapour

4. Classify the following as acidic, basic, neutral ionic, or neutral molecular:
   (a) $\text{KCl}(\text{aq})$ (fertilizer component)
   (b) $\text{HCl}(\text{aq})$ (in stomach)
   (c) sodium hydroxide (oven/drain cleaner)
   (d) ethanol (beverage alcohol)

5. Write the chemical formulas for the following acids:
   (a) aqueous hydrogen chloride (from a gas)
   (b) hydrochloric acid (stomach acid)
   (c) aqueous hydrogen acetate (from a liquid)
   (d) acetic acid (vinegar)
   (e) aqueous hydrogen sulfate (from a liquid)
   (f) sulfuric acid (car battery)
   (g) aqueous hydrogen nitrite (from a gas)
   (h) nitric acid (for making fertilizers)
6. Write accepted names for the following acids:
   (a) $\text{H}_2\text{SO}_3(\text{aq})$ (acid rain)
   (b) $\text{HF}(\text{aq})$ (used for etching glass)
   (c) $\text{H}_2\text{CO}_3(\text{aq})$ (carbonated beverages)
   (d) $\text{H}_2\text{S}(\text{aq})$ (rotten egg odour)
   (e) $\text{H}_3\text{PO}_4(\text{aq})$ (rust remover)
   (f) $\text{HCN}(\text{aq})$ (rat killer)
   (g) $\text{H}_3\text{BO}_3(\text{aq})$ (insecticide)
   (h) $\text{C}_6\text{H}_5\text{COOH}(\text{aq})$ (preservative)

7. Write chemical equations, including the states at SATP, for the following reactions involved in the manufacture and use of sulfuric acid:
   (a) Sulfur reacts with oxygen to produce sulfur dioxide gas.
   (b) Sulfur dioxide reacts with oxygen to produce sulfur trioxide gas.
   (c) Sulfur trioxide gas reacts with water to produce sulfuric acid.
   (d) Sulfuric acid reacts with ammonia gas to produce aqueous ammonium sulfate (a fertilizer).
   (e) Sulfuric acid reacts with rock phosphorus, $\text{Ca}_3(\text{PO}_4)_2(\text{s})$, to produce phosphoric acid and solid calcium sulfate (gypsum).

8. Write chemical equations, including states at SATP, for the following reactions involved in the destructive reactions of acid rain:
   (a) Sulfuric acid in rain reacts with limestone (see Appendix J), causes deterioration of buildings, statues, and gravestones, and produces aqueous hydrogen carbonate (bicarbonate) and solid calcium sulfate.
   (b) Sulfuric acid from rain reacts with solid aluminium silicate in the bottom of a lake and releases aqueous hydrogen silicate (silicic acid) and toxic (to the fish) aqueous aluminium sulfate.

9. Write chemical equations, including states at SATP, for each of the following reactions involved in the control of acid rain:
   (a) Sulfur dioxide emissions can be reduced in the exhaust stack of an oil sands refinery by reacting the sulfur dioxide gas with lime (see Appendix J) and oxygen to produce solid calcium sulfate (gypsum).
   (b) Sulfuric acid in an acid lake can be neutralized by adding slaked lime (see Appendix J) to produce water and solid calcium sulfate.

10. An investigation is planned to explore the conductivity of various categories of substances. Complete the Analysis and Evaluation sections (Parts 2 and 3) of the investigation report.

   **Purpose**
   The purpose of this investigation is to further extend the previously determined empirical definitions of ionic and molecular compounds to include the electrical conductivity of the solid, liquid, and aqueous states of matter.

   **Problem**
   What are the empirical definitions of ionic and molecular compounds?

   **Prediction**
   According to the current definitions of ionic and molecular compounds, ionic compounds are all solids at SATP that form electrically conductive solutions, whereas molecular compounds are solids, liquids, or gases at SATP that form non-conductive solutions.

   **Design**
   Pure samples of water ($\text{H}_2\text{O}$), calcium chloride ($\text{CaCl}_2$), sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$), methanol ($\text{CH}_3\text{OH}$), sodium hydroxide ($\text{NaOH}$), and potassium iodide ($\text{KI}$) are tested for electrical conductivity in the pure state at SATP, in the pure molten state, and in aqueous solution.

   **Evidence**
   Table 5: Electrical Conductivity of Compounds in Different States

<table>
<thead>
<tr>
<th>Chemical formula</th>
<th>Pure state at SATP</th>
<th>Conductivity</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>Pure at SATP</td>
</tr>
<tr>
<td>$\text{H}_2\text{O}$</td>
<td>liquid</td>
<td>none</td>
</tr>
<tr>
<td>$\text{CaCl}_2$</td>
<td>solid</td>
<td>none</td>
</tr>
<tr>
<td>$\text{C}<em>{12}\text{H}</em>{22}\text{O}_{11}$</td>
<td>solid</td>
<td>none</td>
</tr>
<tr>
<td>$\text{CH}_3\text{OH}$</td>
<td>liquid</td>
<td>none</td>
</tr>
<tr>
<td>$\text{NaOH}$</td>
<td>solid</td>
<td>none</td>
</tr>
<tr>
<td>$\text{KI}$</td>
<td>solid</td>
<td>none</td>
</tr>
</tbody>
</table>

*not applicable
Chapter 1 SUMMARY

Outcomes

Knowledge

• classify matter as pure and mixtures as homogeneous and heterogeneous (1.2)
• interpret the periodic table of the elements (1.3)
• use atomic theory to explain the periodic table (1.4)
• classify elements and compounds and know the properties of each class (1.3, 1.4)
• explain and predict chemical formulas for and name ionic and molecular compounds, acids, and bases (1.5, 1.6)
• identify the state of matter of substances (1.5, 1.6)
• write chemical equations when given reactants and products (1.5, 1.6)
• classify scientific knowledge as qualitative and quantitative, as observations and interpretations, and as empirical and theoretical (1.1)

STS

• describe the natures of science and technology (1.1)
• describe the application of some common chemicals (1.3, 1.5, 1.6)

Skills

• use a textbook, a periodic table, and other references efficiently and effectively (1.1–1.6)
• interpret and write laboratory reports (1.1, 1.2, 1.3, 1.4, 1.6)
• select and use diagnostic tests (1.2, 1.3, 1.4, 1.5, 1.6)

Key Terms

1.1 science technology chemistry chemical change chemical reaction physical change observation qualitative observation quantitative observation interpretation empirical knowledge theoretical knowledge empirical hypothesis empirical definition generalization scientific law law of conservation of mass

1.2 matter pure substance mixture heterogeneous mixture homogeneous mixture solution chemical decomposition element entity atom compound chemical formula

1.3 family group period semi-metal

standard ambient temperature and pressure (SATP)
metal nonmetal ductile malleable alkali metal alkaline-earth metal halogen noble gas main group element transition element

1.4 theoretical description theoretical hypothesis theoretical definition theory analogy model mass number atomic number

1.5 ionic compound molecular compound diagnostic test acid base neutral nomenclature aqueous solution formula unit empirical formula

1.6 molecule molecular formula diatomic molecule binary molecular compound

MAKE a summary

1. Prepare a concept map that is centred on pure substances. Include classes of substances along with their properties and nomenclature. See the Key Terms list.

2. Revisit your answers to the Starting Points questions at the start of this chapter. How has your thinking changed?

Go To

The following components are available on the Nelson Web site. Follow the links for Nelson Chemistry Alberta 20–30.

• an interactive Self Quiz for Chapter 1
• additional Diploma Exam-style Review questions
• Illustrated Glossary
• additional IB-related material

There is more information on the Web site wherever you see the Go icon in the chapter.

EXTENSION

Lightning

This video looks into the mystery of lightning from a scientific perspective. An understanding of positive and negative charges, related to the atomic theory, explains the phenomenon.
Chapter 1 REVIEW

Many of these questions are in the style of the Diploma Exam. You will find guidance for writing Diploma Exams in Appendix H. Exam study tips and test-taking suggestions are on the Nelson Web site. Science Directing Words used in Diploma Exams are in bold type.

DO NOT WRITE IN THIS TEXTBOOK.

Part 1

Scientists not only classify natural objects and phenomena, but they also classify knowledge. In both cases, the process of classifying helps them (and us) to better organize knowledge.

1. The classification of knowledge includes
   1. interpretations 2. observations

   Use the above classes of knowledge to classify the following statements, in order.
   • Carbon (as graphite) conducts electricity.
   • Aluminium can pass electrons from atom to atom.
   • Sodium is a metal.
   • Magnesium has two valence electrons.
     ___  ___  ___  ___

2. The classification of knowledge also includes
   1. quantitative observation 2. qualitative observation

   Use the classes of observations to classify the following statements in order:
   • To form an ion, the chlorine atom gains one electron.
   • Gold is malleable.
   • A sodium carbonate solution conducts electricity.
   • The mass of magnesium burned is 2.0 g.
     ___  ___  ___  ___

3. Scientific knowledge can also be classified as
   1. empirical 2. theoretical

   Use the above classes of knowledge to classify the following statements:
   • Molecular compounds form nonconducting aqueous solutions.
   • Ionic compounds dissolve as ions.
   • Acids form conducting aqueous solutions.
   • Bases dissolve to increase the hydroxide ion concentration.
     ___  ___  ___  ___

4. A substance that cannot be decomposed is i. definition for ii.

   The above sentence is completed by the information in which row?
   
   Row i  ii
   A. an empirical a compound
   B. an empirical an element
   C. a theoretical a compound
   D. a theoretical an element

5. The purpose of scientific investigations does not include

   A. creating a concept
   B. verifying a concept
   C. using a concept
   D. testing a concept

6. The criterion that is not used to evaluate a scientific concept is its ability to

   A. explain
   B. predict
   C. describe
   D. prove

7. The relationship of science and technology is best described as

   A. science leading technology
   B. parallel supporting activities
   C. science involving more trial and error
   D. adversarial (in conflict)

Use this information to answer questions 8 and 9.

The format of a laboratory report reflects the basic pattern of scientific research. The following sections of a laboratory report are listed in random order.

1. Analysis 6. Problem
2. Evaluation 7. Procedure
3. Evidence 8. Purpose
5. Materials

8. Once the Purpose and Problem have been chosen, identify, in order, the parts of an investigation report that are done before the work in a laboratory begins.
   ___  ___  ___  ___

9. Identify, in order, the parts of the investigation report that are completed during and after the work in a laboratory.
   ___  ___  ___  ___

10. The Evaluation of an investigation does not involve

    A. evaluating the evidence based upon whether the prediction is verified or falsified
    B. evaluating the evidence based upon whether the design, materials, and procedure are adequate
    C. evaluating the prediction by comparing the answer in the Analysis with the answer in the Prediction
    D. evaluating the hypothesis by comparing the answer in the Analysis with the answer in the Prediction


Science and technology are influential disciplines in our developed country. Understanding the natures of science and technology and their relationship to each other is important to being a citizen of Canada.
11. The name for an entity containing 48 electrons and 48 protons is
   A. cadmium atom
   B. cadmium ion
   C. mercury atom
   D. mercury ion

12. The symbol for an entity containing 80 electrons and 82 protons is
   A. Hg(l)
   B. Pb(s)
   C. Hg^{2+}(aq)
   D. Pb^{2+}(aq)

13. Chemical and Physical Properties of Elements
   1. brittleness
   2. form cations
   3. solid at SATP
   4. good insulators
   5. silvery-grey colour
   6. good conductors of heat
   7. good conductors of electricity

   The four properties shared by cadmium, lead, and mercury, listed in numerical order, are:

14. Identify the molecular compounds in numerical order.

15. Identify the ionic compounds in numerical order.

16. Identify the elements in numerical order.

17. The element that has the least similar properties to the rest is
   A. oxygen
   B. sulfur
   C. bromine
   D. silver

18. The element that does not fit with the chemical properties of the rest is
   A. sodium
   B. potassium
   C. lithium
   D. cerium

19. The melting point of aluminium is __________ °C.

20. Define the following concepts empirically:
   (a) metals
   (b) nonmetals
   (c) molecular compounds
   (d) ionic compounds

21. Define the following concepts theoretically:
   (a) atomic number
   (b) mass number
   (c) isotopes
   (d) anion
   (e) cation

22. Draw energy-level diagrams for the following entities:
   (a) silicon atom
   (b) potassium ion
   (c) fluoride ion
   (d) calcium atom
   (e) sulfide ion

23. Table salt is often used in cooking. Many cook books recommend using salt in the water for cooking green beans. Some books suggest that the salt is necessary for maintaining the green colour of the beans.
   (a) Plan a simple experimental design to test the hypothesis that salt is necessary to maintain the colour of green beans. (Refer to Appendix B.)
   (b) If you know about single and double blind studies (see Appendix B.4), plan a more sophisticated experimental design for this research.
**Evidence**

<table>
<thead>
<tr>
<th>Table 2</th>
<th>Qualitative Analysis Results</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Solution</strong></td>
<td><strong>Conductivity</strong></td>
<td><strong>Litmus paper</strong></td>
</tr>
<tr>
<td>water</td>
<td>none</td>
<td>no change</td>
</tr>
<tr>
<td>1</td>
<td>high</td>
<td>no change</td>
</tr>
<tr>
<td>2</td>
<td>high</td>
<td>blue to red</td>
</tr>
<tr>
<td>3</td>
<td>none</td>
<td>no change</td>
</tr>
<tr>
<td>4</td>
<td>high</td>
<td>red to blue</td>
</tr>
</tbody>
</table>

26. Memorizing is often an initial way of knowing something, such as a chemical formula. What other ways of knowing are available to you?

27. Chemistry is one way of knowing about nature. What are some other ways of knowing about natural phenomena?

**Extension**

28. Use the Internet to investigate how the modern view of fluctuating electrons in an atom compares with an Aboriginal view of a fluctuating universe. Briefly describe your findings.

29. Search for high-technology images, similar to Figure 1, of atoms in elements and molecules, and ions in ionic compounds.

---

**24.** Copy and complete Table 1 by classifying the compounds as ionic or molecular and writing the chemical formulas or IUPAC names.

**Table 1  Ionic and Molecular Compounds**

<table>
<thead>
<tr>
<th>Use</th>
<th>IUPAC name</th>
<th>Ionic or molecular?</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>leavening agent</td>
<td>sodium hydrogen carbonate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>home heating fuel</td>
<td>methane</td>
<td></td>
<td></td>
</tr>
<tr>
<td>bleach</td>
<td></td>
<td></td>
<td>NaClO(s)</td>
</tr>
<tr>
<td>masonry</td>
<td>calcium oxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>dry ice</td>
<td></td>
<td></td>
<td>CO₂(s)</td>
</tr>
<tr>
<td>gas-line antifreeze</td>
<td>methanol</td>
<td></td>
<td></td>
</tr>
<tr>
<td>in laundry detergent</td>
<td>sodium carbonate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>melts ice on</td>
<td>CaCl₂(s)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>sidewalks</td>
<td>C₆H₁₂O₁₁(s)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>fungicide</td>
<td>copper(II) sulfate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>prevents tooth decay</td>
<td></td>
<td></td>
<td>SnF₂(s)</td>
</tr>
<tr>
<td>car batteries</td>
<td>lead(IV) oxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>food seasoning</td>
<td>sodium chloride</td>
<td></td>
<td></td>
</tr>
<tr>
<td>solvent for oils and fats</td>
<td></td>
<td></td>
<td>CC₄(l)</td>
</tr>
<tr>
<td>produces nitric acid</td>
<td></td>
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</tbody>
</table>

25. A qualitative analysis of four compounds is carried out. Complete the Analysis of the following investigation report.

**Purpose**

The purpose of this investigation is to use evidence and empirical definitions to identify four solutions.

**Problem**

Which of the solutions labelled 1, 2, 3, and 4 is KCl(aq), C₂H₅OH(aq), HCl(aq), and Ba(OH)₂(aq)?

**Design**

Each solution is tested with a conductivity apparatus and with litmus paper to determine its identity. A sample of the water used for preparing the solutions is tested for conductivity as a control. Taste tests are ruled out because they are unsafe.

30. Niels Bohr introduced his quantum model of the atom in his 1913 paper. In the last paragraph of the paper, Bohr wrote: “The foundation of the hypothesis has been sought entirely in its relation with Planck’s theory of radiation; ... later it will be attempted to throw some further light on the foundation of it from another point of view.”

The other point of view that Bohr refers to in this paragraph is that of testing the explanatory power of his quantum-model hypothesis by explaining the periodicity of the elements as displayed in the periodic table.

In your own words, use the Bohr theory of the atom to explain the periodic law (the periodicity of the properties of the elements). You can read Bohr’s classic paper on the Nelson Web site.
Chemical reactions are not just something that scientists study in a laboratory. Chemical changes or reactions are an essential part of nature—past, present, and future. For example, lightning initiates many chemical reactions in the atmosphere including the formation of nitrogen compounds that provide nutrients to plants on Earth. Plants carry out many chemical reactions, the most important of which is photosynthesis. Animals eat plants and use the chemical reactions of metabolism to survive and grow. Eventually plants and animals die, and their decomposition is another set of chemical reactions. The cycle of life is a cycle of chemical reactions.

Although you can learn some things about chemicals by observing their physical properties, laboratory work involving chemical reactions reveals a great deal more about the compounds. You can make inferences about chemicals based on the changes that occur in chemical reactions. By studying chemical reactions, you can construct generalizations and laws and, eventually, infer the theoretical structure of the compounds involved.

Initially, theories of the structure of matter attempt to explain the known chemical properties of a substance. The validity of a theory is determined by its ability to both explain and predict changes in matter. How and why do chemicals react? What compounds will form as a result of a reaction? How do we explain the different properties of compounds? Chemical combination represents not only the compounds we know, but also the processes or chemical reactions by which we know them.

**Starting Points**

Answer these questions as best you can with your current knowledge. Then, using the concepts and skills you have learned, you will revise your answers at the end of the chapter.

1. List the kinds of evidence that indicate the occurrence of a chemical reaction.
2. Provide a theoretical explanation as to how and why chemical reactions occur.
3. Describe how an understanding of chemical reaction types can be used to predict the products of a chemical reaction.
4. Describe the process for balancing a chemical reaction equation.
5. The chemical amount of a substance is measured in moles. What is a mole?
Have you ever wondered how magicians set fire to $100 bills without apparently damaging the paper? Or how a server can flambé a dish of food without burning it to a crisp (Figure 1)? Here is a teacher demonstration that might help you answer the puzzle of “how do they do that?”

**Materials:** eye protection, lab apron, protective gloves, a 13 mm test tube with stopper, a small beaker, propan-2-ol (rubbing alcohol), water, table salt, tongs, paper, scissors, safety lighter or match

- Add a few crystals of table salt to a small test tube.
- Add propan-2-ol to nearly half-fill the test tube.
- Add water to nearly fill the test tube.
- Stopper the test tube and invert it several times to dissolve the components.
- Remove the stopper and pour the contents into a small beaker.
- Cut and insert a strip of paper into the alcohol solution.
- Use tongs to remove the paper.
- Remove the beaker of solution from the work area.
- Use the matches or lighter to light the wet paper.
- Dispose of any excess solution in a waste container.
2.1 Science and Technology in Society

The science of chemistry goes hand-in-hand with the technology of chemistry: the skills, processes, and equipment required to make useful products, such as water-resistant adhesives, or to perform useful tasks, such as water purification. Chemists make use of many technologies, from test tubes to computers. Chemists may or may not thoroughly understand a particular chemical technology. For example, the technologies of glass-making and soap-making existed long before scientists could explain these processes. We now use thousands of metals, plastics, ceramics, and composite materials developed by chemical engineers and technologists (Figure 1). However, chemists do not have a complete understanding of superconductors, ceramics, chrome-plating, and some metallurgical processes. Sometimes technology leads science—as in glass-making and soap-making—and sometimes science leads technology. Overall, science and technology complement one another.

Science, technology, and society (STS) are interrelated in complex ways (Figure 2). In this chapter and throughout this book, the nature of science, technology, and STS interrelationships will be introduced gradually, so that you can prepare for decision making about STS issues in the 21st century.

You can acquire specialized knowledge, skills, and attitudes for understanding STS issues by studying science. For example, a discussion of global warming becomes an informed debate when you have specific scientific knowledge about the topic; scientific skills to acquire and test new knowledge; and scientific attitudes and values to guide your thinking and your actions. You also need an understanding of the nature of science and of scientific knowledge.

Scientists have indicated that, at present, both the observations and the interpretations of global warming are inadequate for fully understanding the present phenomenon and for accurately predicting the situation in the future. However, scientists will always state qualifications such as these, even 100 years from now, no matter how much more evidence is available. In a science course, you learn that scientific knowledge is never completely certain or absolute. When scientists testify in courts of law, present reports to parliamentary committees, or publish scientific papers, they tend to avoid authoritarian, exact statements. Instead, they state their results with some degree of uncertainty. In studying science, you learn to look for evidence, to evaluate experiments, and to attach a degree of certainty to scientific statements. You learn to expect and to accept uncertainty, but to search for increasingly greater certainty. This is the nature of scientific inquiry.

Chemicals and chemical processes represent both a benefit and a risk for our planet and its inhabitants. Chemistry has enabled people to produce more food, to dwell more comfortably in homes insulated with fibreglass and polystyrene, and to live longer, thanks to
clean water supplies, more varied diets, and modern drugs. While enjoying these benefits, we also consciously and unconsciously assume certain risks. For example, when chemical wastes are dumped or oil spills into the environment, the effects can be disastrous. Assessing benefits and risks is a part of evaluating advances in science and technology.

The world increasingly depends on science and technology. Our society’s affluence has led to countless technological applications of metal, paper, plastic, glass, wood, and other materials. Thousands of new scientific discoveries and technological advances are made each year. As our society embraces more and more sophisticated technology, we tend to seek technological “fixes” for problems, such as chemotherapy in treating cancer, and the use of fertilizers in agriculture. However, a strictly technological approach to problem solving overlooks the multidimensional nature of the problems confronting us.

Deciding how to use science and technology to benefit society is extremely complex. Most STS issues can be discussed from many different points of view, or perspectives. Even pure scientific research is complicated by economic and social perspectives. For example, should governments increase funding for scientific research when money is needed for social assistance programs? Environmental problems, such as discharge from pulp mills and air pollution, are controversial issues. For rational discussion and acceptable action on STS issues, a variety of perspectives must be taken into account. For example, five of many possible STS perspectives on air pollution are listed below:

- A **scientific perspective** leads to researching and explaining natural phenomena. Research into sources of air pollution and its effects involves a scientific perspective.
- A **technological perspective** is concerned with the development and use of machines, instruments, and processes that have a social purpose. The use of instruments to measure air pollution and the development of processes to prevent air pollution reflect a technological approach to the issue.
- An **ecological perspective** considers relationships between living organisms and the environment. Concern about the effect of a smelter’s sulfur dioxide emissions on plants and animals, including humans, reflects an ecological perspective.
- An **economic perspective** focuses on the production, distribution, and consumption of wealth. The financial costs of preventing air pollution and the cost of repairing damage caused by pollution reflect an economic perspective.
- A **political perspective** involves vote-getting actions and measures. Debate over proposed legislation to control air pollution involves a political perspective.

### Section 2.1 Questions

1. Identify four or more current STS issues.
2. Classify each of the following statements about aluminium as representing a scientific, technological, ecological, economic, or political perspective:
   - (a) Recycled aluminium costs less than one-tenth as much as aluminium produced from ore.
   - (b) Aluminium ore mines in South America have destroyed the natural habitat of plants and animals.
   - (c) Aluminium is refined in Canada using electricity from hydro-electric dams.
   - (d) In Quebec, aluminium is refined using hydro-electric power that some politicians in Newfoundland have claimed belongs to their constituents.
   - (e) In 1886, American chemist Charles Hall discovered through research that aluminium can be produced by using electricity to decompose aluminium oxide dissolved in molten cryolite.
3. Instead of changing their lifestyles, many people look to technology to solve problems caused by the use of technology! Suggest one technological fix and one lifestyle change that would help to solve each of the following problems:
   - (a) Aluminium ore from South America used to produce aluminium metal for beverage cans will be in short supply soon.
   - (b) Pure aluminium cans thrown into the garbage are not magnetic and are, therefore, difficult to separate from the rest of the garbage.
   - (c) People throw garbage into bins for recyclable aluminium cans.
4. Aluminium is used extensively for making beverage cans. List some benefits and risks of this practice. Are there any alternatives that might have equal or better benefits and fewer risks? Be prepared to argue your case.

[DID YOU KNOW?]

**Technological Fixes**

Often our first thought when change is required is to ask for a technology to be invented and/or used to solve the problem. Some examples of technological fixes for societal problems (which themselves may create problems) are:

- escalators and elevators
- cars and airplanes
- radios and televisions
- the Internet and cellphones
- pesticides and fertilizers
- gasoline and plastics

[LEARNING TIP]

A mnemonic that may be used to recall these five STS perspectives is STEEP.
2.2 Changes in Matter

The explanation of natural events is one of the aims in science. Careful observation, leading to the formation of a concept or theory, and followed by testing and evaluating the ideas involved, defines the basic process scientists use to increase understanding of the changes going on in the world around us. A useful way to begin is to classify the types of changes that occur in matter. Changes in matter can be explained at three levels according to size. Modern scientists study and discuss matter at a macroscopic (naked eye observable) level, or at a microscopic (too small to see without a microscope) level, or at a molecular (smallest entities of a substance) level. To understand their observations at a molecular level, chemists usually start by basing their explanations on the atomic theory proposed by John Dalton in 1803.

Types of Changes in Matter

Chemists often describe changes in matter as a physical change, chemical change, or nuclear change (Figure 1), depending on whether they believe that a change has occurred in the molecules, electrons, or nuclei of the substance being changed. The quantity of energy associated with every change in matter can also help classify the type of change.

Physical changes are any changes where the fundamental entities remain unchanged at a molecular level, such as the phase changes of evaporation and melting. There is no change in the written formula of the substance involved (Figure 1(a)). Dissolving a chemical is usually classified as a physical change. Other examples include changes in physical structure that change appearance, like grinding a shiny piece of copper into a fine powder that is black in colour, but is still copper metal. Physical changes in matter usually involve relatively small amounts of energy change.

A chemical change involves some kind of change in the chemical bonds within the fundamental entities (between atoms and/or ions) of a substance, and is represented by a change in the written formula (Figure 1(b)). At least one new substance is formed, with physical and chemical properties different from those of the original matter. Normally, chemical changes involve larger energy changes than physical changes.

Nuclear changes (changes within the nucleus) create entirely new atomic entities (Figure 1(c)). These entities are represented by formulas that show new atomic symbols, different from those of the original matter. Nuclear changes involve extremely large changes in energy, which allow them to be identified. In 1896, Henri Becquerel noticed the continuous production of energy from a piece of rock that showed no other changes at all. His observation led to the discovery of radioactivity—a nuclear change.

Physical, chemical, and nuclear changes can be described both empirically and theoretically. Table 1 provides these descriptions, along with an example of each. In this course, we will focus our attention almost entirely on chemical change. We can use classification systems, such as Table 1, to help us to organize our knowledge.

Figure 1
(a) Hydrogen is liquified at −253 °C. H₂(g) → H₂(l)
(b) Hydrogen is burning—as it does during the space shuttle launch and in hydrogen-fueled automobiles. 2H₂(g) + O₂(g) → 2H₂O(g)
(c) Hydrogen is undergoing nuclear fusion on the Sun and is being converted into helium. H(g) + H(g) → He(g)

WEB Activity
Case Study—States of Matter and Changes in Matter

Watch the movie about properties of matter and physical changes.

(a) What properties of solids, liquids, and gases make them different from one another?
(b) What kinds of physical changes are not mentioned in the movie?
(c) What are some clues that a chemical reaction has taken place?
The Kinetic Molecular Theory

Scientists observed gas pressure, diffusion, and chemical reactions, and eventually explained their observations using the concept of molecular motion. The idea of molecular motion led to the kinetic molecular theory, which has become a cornerstone of modern science.

The central idea of the kinetic molecular theory is that the smallest entities of a substance are in continuous motion. These entities may be atoms, ions, or molecules. As they move about, the entities collide with each other and with objects in their path (Figure 2).

How and Why Chemical Reactions Occur

Chemical changes are also called chemical reactions. To explain chemical reactions, we can expand the kinetic molecular theory to create a theory of chemical reactions. According to the kinetic molecular theory, the entities of a substance are in continuous, random motion. This motion inevitably results in collisions between the entities. If different substances are present, all the different entities will collide randomly with each other. If the collision has a certain orientation and sufficient energy, the components of the entities will rearrange to form new entities. The rearrangement of entities that occurs is the chemical reaction. This general view of a chemical reaction is known as the collision–reaction theory (Figure 3).

Table 1 Physical, Chemical, and Nuclear Change

<table>
<thead>
<tr>
<th>Change</th>
<th>Empirical description</th>
<th>Theoretical description</th>
</tr>
</thead>
<tbody>
<tr>
<td>physical</td>
<td>• state or energy change&lt;br&gt;• solid → liquid → gas&lt;br&gt;• no new substance&lt;br&gt;• small energy change</td>
<td>• $H_2(g) \rightarrow H_2(l) +$ energy&lt;br&gt;• $H_2O(s) \rightarrow H_2O(l) \rightarrow H_2O(g)$&lt;br&gt;• no new molecules&lt;br&gt;• intermolecular forces broken and made</td>
</tr>
<tr>
<td>chemical</td>
<td>• colour, odour, state, and/or energy change&lt;br&gt;• new substance formed&lt;br&gt;• new permanent properties&lt;br&gt;• medium energy change</td>
<td>• $2H_2(g) + O_2(g) \rightarrow 2H_2O(g) +$ energy&lt;br&gt;• old (reactants) → new (products)&lt;br&gt;• atoms/ions/electrons rearranged&lt;br&gt;• chemical bonds broken and made</td>
</tr>
<tr>
<td>nuclear</td>
<td>• often radiation emitted&lt;br&gt;• new elements formed&lt;br&gt;• enormous energy change</td>
<td>• $^1H + ^1H \rightarrow ^3He +$ energy&lt;br&gt;• new atoms formed&lt;br&gt;• nuclear bonds broken and made</td>
</tr>
</tbody>
</table>

The collision–reaction theory explains that chemical entities must collide with the correct orientation to react.
A theoretical explanation of why chemical reactions occur is partially covered in Chapter 1. Atoms often react in order to obtain a more stable electron arrangement (often an octet of electrons like the nearest noble gas). The how and why of chemical reactions can be summarized in a statement such as, “when chemical entities collide, they may exchange or share electrons to obtain a more favourable (stable) electron arrangement.”

**Chemical Reactions**

Recall that chemical reactions produce new substances. How do you know if an unfamiliar change is a chemical reaction? Certain characteristic evidence is associated with chemical reactions (Table 2).

<table>
<thead>
<tr>
<th>Evidence</th>
<th>Description and example</th>
</tr>
</thead>
<tbody>
<tr>
<td>colour change</td>
<td>The final product(s) may have a different colour than the colour(s) of the starting material(s). For example, the solution changes from colourless to blue.</td>
</tr>
<tr>
<td>odour change</td>
<td>The final material(s) may have a different odour than the odour(s) of the starting material(s). For example, mixing solutions of sodium acetate and hydrochloric acid produces a mixture that smells like vinegar.</td>
</tr>
<tr>
<td>state change</td>
<td>The final material(s) may include a substance in a state that differs from the starting material(s). Most commonly, either a gas or a solid (precipitate) is produced.</td>
</tr>
<tr>
<td>energy change</td>
<td>When a chemical reaction occurs, energy in the form of heat, light, sound, or electricity is absorbed or released. For most chemical reactions, the energy absorbed or released is in the form of heat. A common example of an energy change is the combustion, or burning, of a fuel. If energy is absorbed, the reaction is endothermic. If energy is released, the reaction is exothermic.</td>
</tr>
</tbody>
</table>

A diagnostic test is a short and specific laboratory procedure, with expected evidence and analysis, that is used as an empirical test for the presence of a substance. Diagnostic tests increase the certainty that a new substance has formed in a chemical reaction. Appendix C.3 describes diagnostic tests for chemicals such as hydrogen and oxygen. If the diagnostic test entails a single step for a specific chemical, you may find it convenient to summarize this test using the format, “If [procedure] and [evidence], then [analysis].” An example of a diagnostic test is shown in Figure 4.

**Conservation of Mass in Chemical Changes**

Experimenters have found that the total mass of matter present after a chemical change is always the same as the total mass present before the change, no matter how different the new substances appear. This finding is called the law of conservation of mass. The law of conservation of mass was one of the compelling reasons why scientists accepted the atomic theory of matter. If a chemical change is thought of as a rearrangement of entities at the molecular level, then it is simple to argue that the mass must be constant. The individual entities do not change, except in the ways they are associated with each other.

**Communicating Chemical Reactions**

A balanced chemical equation is one in which the total number of each kind of atom or ion in the reactants is equal to the total number of the same kind of atom or ion in the products.
Figure 5 shows the balanced chemical equation and molecular models representing the reaction of nitrogen dioxide gas and water to produce nitric acid and nitrogen monoxide gas. By studying the molecular models, you can see there are three nitrogen atoms on the reactant side and three on the product side of the equation arrow. Likewise, there are seven oxygen atoms on both sides of the equation arrow, and two hydrogen atoms.

If more than one molecule is involved (for example, three molecules of nitrogen dioxide), then a number called a coefficient is placed in front of the chemical formula. In this example, three molecules of nitrogen dioxide and one molecule of water react to produce two molecules of nitric acid and one molecule of nitrogen monoxide. Coefficients are part of a balanced chemical equation and should not be confused with formula subscripts, which are part of the chemical formula for a substance.

A substance’s state of matter is given in parentheses after the chemical formula. It is not part of the theoretical description given by the molecular models. Chemical formulas showing states of matter provide both a theoretical and an empirical description of a substance.

**SUMMARY**

**Chemical Reaction Equations**

- A chemical reaction is communicated by a balanced chemical equation in which the same number of each kind of atom or ion appears on the reactant and product sides of the equation.
- A coefficient in front of a chemical formula in a chemical equation communicates the number of molecules or formula units of a reactant or product that are involved in the reaction.
- Within formulas, a numerical subscript communicates the number of atoms or ions present in one molecule or formula unit of a substance.
- A state of matter in parentheses in a chemical equation communicates the physical state of the reactants and products at SATP.
### Section 2.2 Questions

1. Provide two examples each of physical, chemical, and nuclear changes.

2. What different entities are rearranged during physical, chemical, and nuclear changes?

3. According to the collision–reaction theory, identify the requirements for a chemical reaction to take place.

4. What is the purpose of classification systems, such as those for types of changes by substances?

5. List four changes that can be used as evidence for chemical reactions.

6. Provide two examples from everyday life of each of the four types of changes listed in your answer to the previous question.

7. Use an “If [procedure] and [evidence], then [analysis]” format to write diagnostic tests for an acid and for hydrogen (see Appendix C.4).

8. Identify one scientific law that led John Dalton to create the theory that atoms are conserved in a chemical reaction.

9. Distinguish between a formula subscript, such as H₂, and a coefficient, such as 2 H. How are they similar and how are they different?

10. An investigation is conducted to observe and classify evidence of chemical changes. The substances are mixed and the evidence obtained is recorded in Table 3. (a) Match as many mixtures as possible to each of the four categories of evidence of chemical reactions (see Table 2, page 48). (b) Which mixture did not appear to have a chemical reaction? How certain are you about this interpretation? (c) In general terms, what additional laboratory work could be done to improve the certainty that a chemical reaction has occurred?

### Table 3 Evidence for Chemical Reactions

<table>
<thead>
<tr>
<th>Mixture</th>
<th>Procedure</th>
<th>Evidence</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>A zinc strip is briefly dipped into a hydrochloric acid solution.</td>
<td>Colourless gas bubbles formed on the strip.</td>
</tr>
<tr>
<td>2</td>
<td>A couple of drops of blue bromothymol blue solution are added to hydrochloric acid.</td>
<td>The blue bromothymol blue solution turned yellow.</td>
</tr>
<tr>
<td>3</td>
<td>A few drops of silver nitrate solution are added to hydrochloric acid.</td>
<td>A white solid (precipitate) formed.</td>
</tr>
<tr>
<td>4</td>
<td>Hydrochloric acid is added to a sodium acetate solution.</td>
<td>The odour of vinegar was evident.</td>
</tr>
<tr>
<td>5</td>
<td>Ammonium nitrate crystals are stirred into water.</td>
<td>The water (solution) felt cool.</td>
</tr>
<tr>
<td>6</td>
<td>Hydrochloric acid is added to a sodium bicarbonate solution.</td>
<td>Colourless gas bubbles formed.</td>
</tr>
<tr>
<td>7</td>
<td>A couple of drops of phenolphthalein solution are added to an ammonia solution.</td>
<td>The colourless phenolphthalein turned red.</td>
</tr>
<tr>
<td>8</td>
<td>Sodium hydroxide solution is added to a cobalt(II) chloride solution.</td>
<td>A blue and/or pink solid (precipitate) formed.</td>
</tr>
<tr>
<td>9</td>
<td>Sodium nitrate solution is added to a potassium chloride solution.</td>
<td>No change was observed.</td>
</tr>
<tr>
<td>10</td>
<td>A copper wire is placed into a silver nitrate solution.</td>
<td>Silvery crystals formed on the wire.</td>
</tr>
</tbody>
</table>

### Extension

11. A common media mistake is to refer to the dissolving of a chemical in water or the reaction of a metal with a solution as melting. How would you explain to the media that their concept of melting is incorrect?

12. There is debate among chemists as to whether dissolving is a physical change or a chemical change. What does the existence of a debate tell you about classification systems?

13. Evidence-based reasoning is a mainstay of scientific work. There are a few terms that a scientifically literate person needs to know to read newspapers, magazines, and scientific reports (see Appendix B.4.) What do the following terms mean with respect to scientific research? (a) anecdotal evidence (b) sample size (c) replication (d) placebo

14. Scholars publish their research in a refereed (peer-reviewed) journal. These journals are like scholarly magazines. Scientists who submit their articles for publication have their research reviewed by their peers (experts in their field of study). Many research reports submitted for publication are rejected because of some fault in the study. As a referee (peer-reviewer), critique the following experimental designs. (a) One group of 10 patients is given experimental medication for an illness while another group of 10 patients is not given the medication. The health of all twenty patients is monitored for six months. (b) One chemistry teacher completes a unit of study without doing any laboratory work; another teacher completes the same unit of study by doing four laboratory investigations. Student achievement is compared on the results of a unit test.
You are already familiar with some terms used to define convenient numbers (Table 1). For example, a dozen is a convenient number referring to such items as eggs or doughnuts. Since atoms, ions, and molecules are extremely small entities, a convenient number for them must be much greater than a dozen. A convenient amount of substance—also called the chemical amount—is the SI quantity for the number of entities in a substance. It is measured in units of moles (SI symbol, mol). Modern methods of estimating this number of entities have led to the value $6.02 \times 10^{23}$. This value is called Avogadro’s number, named after the Italian chemist, Amedeo Avogadro (1776–1856). (Avogadro did not determine the number, but he created the research idea.) A mole is the unit of chemical amount of substance with the number of entities corresponding to Avogadro’s number. For example:

- one mole of sodium is $6.02 \times 10^{23}$ Na atoms
- one mole of chlorine is $6.02 \times 10^{23}$ Cl$_2$ molecules
- one mole of sodium chloride is $6.02 \times 10^{23}$ NaCl formula units

Essentially, a mole represents a number ($6.02 \times 10^{23}$, Avogadro’s number), just as a dozen represents the number 12.

Although the mole represents an extraordinarily large number, a mole of a substance is an observable quantity that is convenient to measure and handle. Figure 1 shows a mole of each of three common substances: one element, one ionic compound, and one molecular compound. In each case, a mole of entities is a sample size that is convenient for lab work.

### Translating Balanced Chemical Equations
A balanced chemical equation can be interpreted theoretically in terms of individual atoms, ions, or molecules, or groups of them. Consider the reaction equation for the industrial production of the fertilizer ammonia:

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

- 1 molecule
- 3 molecules
- 2 molecules
- 1 dozen molecules
- 3 dozen molecules
- 2 dozen molecules
- 1 mol nitrogen
- 3 mol hydrogen
- 2 mol ammonia
- $6.02 \times 10^{23}$ molecules
- $3(6.02 \times 10^{23})$ molecules
- $2(6.02 \times 10^{23})$ molecules
- $6.02 \times 10^{23}$ molecules
- $18.06 \times 10^{23}$ molecules
- $12.04 \times 10^{23}$ molecules

Note that the numbers in each row are in the same ratio (1:3:2) whether individual molecules, large numbers of molecules, or moles are considered. When moles are used to express the coefficients in the balanced equation, the ratio of reacting amounts is called the mole ratio.

A complete translation of the balanced chemical equation for the formation of ammonia is: “One mole of nitrogen gas and three moles of hydrogen gas react to form two moles of ammonia gas.” This translation includes all the symbols in the equation, including coefficients and states of matter.

### Table 1 Convenient Numbers

<table>
<thead>
<tr>
<th>Quantity</th>
<th>Number</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>pair</td>
<td>2</td>
<td>shoes</td>
</tr>
<tr>
<td>dozen</td>
<td>12</td>
<td>eggs</td>
</tr>
<tr>
<td>gross</td>
<td>144</td>
<td>pencils</td>
</tr>
<tr>
<td>ream</td>
<td>500</td>
<td>paper</td>
</tr>
<tr>
<td>mole</td>
<td>$6.02 \times 10^{23}$</td>
<td>molecules</td>
</tr>
</tbody>
</table>

---

**Figure 1** These amounts of carbon, table salt, and sugar each contain about a mole of entities (atoms, formula units, molecules) of the substance. The mole represents a convenient and specific quantity of a chemical. Note that equal numbers of entities can have very different volumes and masses.
Balancing Chemical Equations

A chemical equation is a simple, precise, logical, and international method of communicating the experimental evidence of a reaction. The evidence used when writing a chemical equation is often obtained in stages. First there are some general observations that suggest a chemical change has occurred. These are likely followed by a series of diagnostic tests to identify the products of the reaction. At this stage an unbalanced chemical equation can be written, and then the theory of conservation of atoms can be used to predict the coefficients necessary to balance the reaction equation. In most cases, trial and error, as well as intuition and experience, play an important role in successfully balancing chemical equations. The following summary outlines a systematic approach to balancing equations. Use it as a guide as you study Sample Problem 2.1.

**COMMUNICATION example**

Translate the following chemical equation into an English sentence.

$$6 \text{ CO}_2(g) + 6 \text{ H}_2\text{O}(l) \rightarrow C_6\text{H}_{12}\text{O}_6(aq) + 6 \text{ O}_2(g)$$

**Solution**

Six moles of carbon dioxide gas react with six moles of liquid water to produce one mole of aqueous glucose and six moles of oxygen gas.

**Balancing Chemical Equations**

Step 1: Write the chemical formula for each reactant and product, including the state of matter for each one.

Step 2: Try balancing the atom or ion present in the greatest number. Find the lowest common multiple to obtain coefficients to balance this particular atom or ion.

Step 3: Repeat step 2 to balance each of the remaining atoms and ions.

Step 4: Check the final reaction equation to ensure that all atoms and ions are balanced.

**SAMPLE problem 2.1**

A simple technology for recycling silver is to trickle waste solutions containing silver ions over scrap copper. Copper metal reacts with aqueous silver nitrate to produce silver metal and aqueous copper(II) nitrate (Figure 2). Write the balanced chemical equation.

Step 1: $\text{Cu}(s) + ? \text{AgNO}_3(aq) \rightarrow ? \text{Ag}(s) + ? \text{Cu(NO}_3)_2(aq)$

Step 2: Oxygen atoms are present in the greatest number as part of the nitrate ion, so balance this first. Balance the nitrate ion as a group.

$\text{Cu}(s) + 2 \text{AgNO}_3(aq) \rightarrow \text{Ag}(s) + \text{Cu(NO}_3)_2(aq)$

Step 3: Balance Ag and Cu atoms. (Always balance elements last.)

$\text{Cu}(s) + 2 \text{AgNO}_3(aq) \rightarrow \text{2Ag}(s) + \text{Cu(NO}_3)_2(aq)$

Step 4: The chemical amounts in moles of copper, silver, and nitrate are one, two, and two on both the reactant and the product sides of the equation arrow. (This is a mental check; no statement is required.)

Translation: One mole of solid copper reacts with two moles of aqueous silver nitrate to produce two moles of solid silver and one mole of aqueous copper(II) nitrate.
Use the following techniques for balancing chemical equations:

- Persevere and realize that, like solving puzzles, several attempts may be necessary for more complicated chemical equations.
- The most common student error is to use incorrect chemical formulas to balance the chemical equation. Always write correct chemical formulas first and then balance the equation as a separate step.
- If polyatomic ions remain intact, balance them as a single unit.
- Delay balancing any atom that is present in more than two substances in the chemical equation until all other atoms or ions are balanced. (Oxygen atoms in several entities is a common example.)
- Balance elements (entities with only one kind of atom) last.
- If a fractional coefficient is required to balance an atom, multiply all coefficients by the denominator of the fraction to obtain integer values. (Balancing with fractions is correct, but not preferred.)

For example, in balancing the following reaction equation, hydrogen atoms are balanced first, then nitrogen, and oxygen is balanced last. This equation requires 7 mol of oxygen atoms:

\[ 2\text{NH}_3(g) + ?\text{O}_2(g) \rightarrow 3\text{H}_2\text{O}(g) + 2\text{NO}_2(g) \]

The only number that can balance the oxygen atoms is \( \frac{7}{2} \). By doubling all coefficients, the reaction equation can then be balanced using only integers:

\[ 4\text{NH}_3(g) + 7\text{O}_2(g) \rightarrow 6\text{H}_2\text{O}(g) + 4\text{NO}_2(g) \]

---

**Section 2.3 Questions**

1. Translate the following English sentences into internationally understood balanced chemical equations:
   (a) Two moles of solid aluminium and three moles of aqueous copper(II) chloride react to form three moles of solid copper and two moles of aqueous aluminium chloride. (This reaction does not always produce the expected products listed here.)
   (b) One mole of solid copper reacts with two moles of hydrochloric acid to produce one mole of hydrogen gas and one mole of copper(II) chloride. (When tested in the laboratory, this prediction of products is falsified.)
   (c) Two moles of solid mercury(II) oxide decomposes to produce two moles of liquid mercury and one mole of oxygen gas. (This decomposition reaction is a historical but dangerous method of producing oxygen. Research an MSDS for mercury(II) oxide.)
   (d) Methanol (used in windshield washer antifreeze and as a fuel) is produced from natural gas in world-scale quantities in Medicine Hat, Alberta by the following reaction series:
      (i) One mole of methane gas reacts with one mole of steam to produce one mole of carbon monoxide gas and three moles of hydrogen gas.
      (ii) One mole of carbon monoxide gas reacts with two moles of hydrogen gas to produce one mole of liquid methanol.

2. Translate each of the following chemical equations into an English sentence including the chemical amounts and states of matter for all the substances involved:
   (a) Fire-starters for camp fires often involve the following reaction. Methanol is also a fondue fuel.
      \[ 2\text{CH}_3\text{OH}(l) + 3 \text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 4 \text{H}_2\text{O}(g) \]
   (b) Phosphoric acid for fertilizer production is produced from rock phosphorus at Fort Saskatchewan, Alberta:
      \[ \text{Ca}_3(\text{PO}_4)_2(s) + 3 \text{H}_2\text{SO}_4(aq) \rightarrow 2\text{H}_3\text{PO}_4(aq) + 3 \text{CaSO}_4(s) \]
   (c) The reaction of sodium with water is potentially dangerous.
      \[ 2\text{Na}(s) + 2 \text{H}_2\text{O}(l) \rightarrow \text{H}_2(g) + 2\text{NaOH}(aq) \]
   (d) Sulfuric acid can be used as a catalyst to dehydrate sugar.
      \[ \text{C}_12\text{H}_22\text{O}_{11}(s) \rightarrow 12 \text{C}(s) + 11 \text{H}_2\text{O}(g) \]

---

**Learning Tip**

When water is a product of a reaction, it is often produced as a gas (vapour), \( \text{H}_2\text{O}(g) \). The heat of combustion produces temperatures above the boiling point of water, 100 °C. If the temperature is low and/or the humidity is high, the water condenses to \( \text{H}_2\text{O}(l) \) or freezes to \( \text{H}_2\text{O}(s) \), as seen in vapour trails of cars and jets.
3. Which of the following chemical equations is balanced correctly?
(a) \( \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g}) \)
(b) \( 2 \text{NaOH}(\text{aq}) + \text{Cu(ClO}_3\text{)}_2(\text{aq}) \rightarrow \text{Cu(OH)}_2(\text{s}) + 2 \text{NaClO}_3(\text{aq}) \)
(c) \( \text{Pb}(\text{s}) + \text{AgNO}_3(\text{aq}) \rightarrow \text{Ag}(\text{s}) + \text{Pb(NO}_3\text{)}_2(\text{aq}) \)
(d) \( 2 \text{NaHCO}_3(\text{s}) \rightarrow \text{Na}_2\text{CO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \)

4. Write a balanced chemical equation for each of the following reactions. Assume that substances are pure and at SATP unless the states of matter are given. Also classify the primary perspective presented in the accompanying statements: The possible perspectives are scientific, technological, ecological, economic, and political.
(a) Research indicates that sulfur dioxide gas reacts with oxygen in the air to produce sulfur trioxide gas.
(b) Sulfur trioxide gas travelling across international boundaries causes disagreements between governments.
(c) The means exist for industry to reduce sulfur dioxide emissions; for example, by treatment with lime.
(d) Restoring acidic lakes to normal pH (acid–base balance) is expensive; for example, adding lime to lakes using aircraft (Figure 3).
(e) Fish in overly acidic lakes may die from mineral poisoning due to the leaching of, for example, aluminium ions from lake bottoms.
(f) Chemical engineers prepare sodium fluoride in large batches for city water or toothpaste distributors.

5. Balance the following equations that communicate reactions that occur before, during, and after the formation of acid rain.
(a) \( \text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) \)
(b) \( \text{S}_8(\text{s}) + 4 \text{O}_2(\text{g}) \rightarrow 8 \text{SO}_2(\text{g}) \)
(c) \( \text{CaSiO}_3(\text{s}) + 2 \text{H}_2\text{SO}_3(\text{aq}) \rightarrow \text{H}_2\text{SiO}_3(\text{aq}) + \text{CaSO}_3(\text{s}) \)
(d) \( \text{CaCO}_3(\text{s}) + 2 \text{HNO}_3(\text{aq}) \rightarrow \text{H}_2\text{CO}_3(\text{aq}) + \text{Ca(NO}_3\text{)}_2(\text{aq}) \)
(e) \( \text{Al}(\text{s}) + 3 \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Al}_2(\text{SO}_4\text{)}_3(\text{aq}) + 3 \text{H}_2(\text{g}) \)
(f) \( \text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_3(\text{aq}) \)
(g) \( \text{Fe}(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{H}_2(\text{g}) + \text{Fe}_2(\text{SO}_3\text{)}_3(\text{aq}) \)
(h) \( \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{CO}_3(\text{aq}) \)
(i) \( \text{CH}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \)
(j) \( \text{FeS}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{FeO}(\text{s}) + \text{SO}_2(\text{g}) \)
(k) \( \text{H}_2\text{S}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g}) + \text{SO}_2(\text{g}) \)
(f) \( \text{CaCO}_3(\text{s}) + 2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CaSO}_4(\text{s}) + \text{CO}_2(\text{g}) \)

6. Balance the following chemical reaction equations and draw molecular models to represent the number and kind of atoms in each molecule.
(a) Hydrogen is used as fuel for the space shuttle.
(b) Ammonia fertilizer is produced for agricultural use.
(c) Hydrogen chloride gas is produced and then dissolved to make hydrochloric acid to etch concrete.

Extension
7. Scientists have a high standard for accepting knowledge as valid/acceptable. Write a short procedure for a double blind experimental design to test the effectiveness of using lime to neutralize acid lakes. Reference Appendix B.4.
Chemists created the concept of the mole. They needed a convenient unit for determining the quantities of chemicals that react and/or are produced in a chemical reaction. You have seen that the mole is a useful unit for communicating reacting amounts in a balanced chemical equation. Now you will learn how to measure chemical amount (in moles) indirectly, by measuring mass directly.

**Molar Mass**

The molar mass, \( M \), of a substance is the mass of one mole of the substance. The unit for molar mass is grams per mole (g/mol) (Figure 1). Each substance has a different molar mass, which can be calculated as follows:

1. Write the correct chemical formula for the substance.
2. Determine the chemical amount (in moles) of each atom (or monatomic ion) in one mole of the chemical.
3. Use the atomic molar masses from the periodic table and the chemical amounts (in moles) to determine the mass of one mole of the chemical.
4. Communicate the mass of one mole as the molar mass in units of grams per mole, precise to two decimal places; for example, 78.50 g/mol.

You may also think of the molar mass as a ratio of the mass of a particular chemical to the amount of the chemical in moles. Molar mass is a convenient factor to use when converting between mass and chemical amount. In Sample Problems 2.2 and 2.3, \( M \) represents the molar mass, and the numbers and atomic molar masses are written out.

When doing calculations involving molar masses, remember that the relevant SI quantities and units are:

- mass, \( m \), in grams, g
- chemical amount, \( n \), in moles, mol
- molar mass, \( M \), in grams per mole, g/mol

**SAMPLE problem 2.2**

Determine the molar mass of water.

**Step 1:** Write the correct chemical formula for water: \( H_2O(l) \).

**Step 2:** Determine the chemical amount of each atom in one mole of \( H_2O(l) \): one mole of \( H_2O \) is composed of 2 mol \( H \) + 1 mol \( O \).

**Step 3:** Find the mass of one mole of water using the number of atoms and their atomic molar masses from the periodic table.

\[
\begin{align*}
m_{H_2O} &= 2 \times M_H + 1 \times M_O \\
&= 2 \text{ mol } \times 1.01 \text{ g/mol } + 1 \text{ mol } \times 16.00 \text{ g/mol} \\
&= 18.02 \text{ g}
\end{align*}
\]

**Step 4:** Since this quantity is the mass of one mole of water, the molar mass of water is 18.02 g/mol; that is,

\[
M_{H_2O} = 18.02 \text{ g/mol}
\]

---

**Learning Tip**

When adding and subtracting measured values, the answer should have the same precision (number of decimal places) as the value with the least precision (number of decimal places). See Appendix F.3.

---

**DID YOU KNOW?**

**SI**

SI is the Système International d’Unités or International System of Units. SI is used in all languages to denote this system of units. SI was created by international agreement in 1960. Canada officially became an SI metric country in the early 1970s.
Mass–Amount Conversions
In order to use the mole ratio from the balanced equation to determine the masses of reactants and products in chemical reactions, you must be able to convert a mass to a chemical amount (an amount in moles) and vice versa. To do this, you use either the molar mass as a conversion factor in grams per mole (g/mol), or the reciprocal of the molar mass, in moles per gram (mol/g), and cancel the units (Figure 2). Examples of each conversion follow; \( n \) represents chemical amount in moles and \( m \) represents mass in grams.

**COMMUNICATION example 1**

Calcium carbonate helps to neutralize acidic soil under spruce trees. Convert a mass of 1500 g of calcium carbonate to a chemical amount.

**Solution**

\[
M_{CaCO_3} = 100.09 \text{ g/mol} \\

n_{CaCO_3} = 1500 \text{ g} \times \frac{1 \text{ mol}}{100.09 \text{ g}} = 14.99 \text{ mol}
\]

In Communication Example 1, the appropriate conversion factor is the reciprocal of the molar mass. The certainty of the value for chemical amount (14.99 mol) is four significant digits, resulting from the least certain of four significant digits for 1500 g and five significant digits for 100.09 g/mol.
Section 2.4

Sodium sulfate is mined from lakes and deposits along the Alberta–Saskatchewan border. Convert a reacting amount of 3.46 mmol of sodium sulfate into mass in grams.

Solution

\[ M_{\text{Na}_2\text{SO}_4} = 142.04 \text{ g/mol} \]

\[ m_{\text{Na}_2\text{SO}_4} = 3.46 \text{ mmol} \times \frac{142.04 \text{ g}}{1 \text{ mmol}} = 491 \text{ mg} \]

or

\[ m_{\text{Na}_2\text{SO}_4} = 3.46 \text{ mmol} \times \frac{1 \text{ g}}{1000 \text{ mmol}} \times \frac{142.04 \text{ g}}{1 \text{ mol}} = 0.491 \text{ g} \]

In Communication Example 2, the certainty of the value for mass (491 mg or 0.491 g) is three significant digits, resulting from the least certain of 3.46 mmol and 142.04 g/mol.

Section 2.4 Questions

It helps to memorize the chemical formula or name for each technological or natural substance marked with an asterisk (*).

1. Calculate the molar mass of each of the following substances. The molar masses of water and of carbon dioxide should be memorized for efficient work:
   (a) \( \text{H}_2\text{O}(l) \) (water)*
   (b) \( \text{CO}_2(g) \) (respiration product)*
   (c) \( \text{NaCl}(s) \) (pickling salt, sodium chloride)*
   (d) \( \text{C}_12\text{H}_22\text{O}_11(s) \) (table sugar, sucrose)*
   (e) \( (\text{NH}_4)\text{2Cr}_2\text{O}_7(s) \) (ammonium dichromate)

2. Communicate the certainty of the following measured or calculated values as a number of significant digits.
   (a) 16.05 g
   (b) 7.0 mL
   (c) 10 cm²
   (d) 0.563 kg
   (e) 0.000 5 L
   (f) 90.00 g/mol

3. Perform the following calculations and express the answer to the correct certainty (number of significant digits).
   (a) \( n_{\text{Cu}} = 7.46 \text{ g} \times \frac{1 \text{ mol}}{63.55 \text{ g}} = \)
   (b) \( n_{\text{C}} = 2.0 \text{ mol} \times \frac{12.01 \text{ g}}{1 \text{ mol}} = \)
   (c) \( n_{\text{Na}_2\text{SO}_4} = 100.0 \text{ mL} \times \frac{0.500 \text{ mol}}{1 \text{ L}} = \)
   (d) \( n_{\text{C}_2\text{H}_5\text{OH}} = 0.0500 \text{ mol} \times \frac{24.97 \text{ g}}{1 \text{ mol}} = \)

4. Perform the following more advanced calculations by using the precision and/or uncertainty rule where appropriate.
   (a) \( m_{\text{NH}_3} = 101 \text{ mol} \times \frac{17.04 \text{ g}}{1 \text{ mol}} = \)
   (b) \( n_{\text{CuSO}_4} = 250.0 \text{ mol} \times \frac{1 \text{ L}}{5.00 \text{ mol}} = \)

5. Calculate the chemical amount of pure substance present in each of the following samples:
   (a) 40.0 g of propane, \( \text{C}_3\text{H}_8(l) \), in a camp stove cylinder
   (b) A 500 g box of pickling salt*
   (c) A 10.00 kg bag of table sugar*
   (d) 325 mg of acetylsalicylic acid (ASA), \( \text{C}_9\text{H}_8\text{O}_4\text{COOH}(s) \), in a headache relief tablet
   (e) 150 g of isopropanol (rubbing alcohol), \( \text{CH}_3\text{CH}_2\text{OH}(l) \), from a pharmacy

6. Calculate the mass of each of the following specified chemical amounts of pure substances:
   (a) 4.22 mol of ammonia in a window-cleaning solution*
   (b) 0.224 mol of sodium hydroxide (lye) in a drain-cleaning solution*
   (c) \( 15.5 \text{ mmol} = \frac{10.00 \text{ mL}}{2} = \)
   (d) \( V_{\text{avg}} = \frac{13.6 \text{ mL} + 13.5 \text{ mL} + 13.6 \text{ mL}}{3} = \)
   (e) \( \% \text{ difference} = \frac{|3.67 g - 3.61 g|}{3.61 g} \times 100 = \)
   (f) \( Q = 50.0 \text{ g} \times 4.19 \text{ J/(g}°\text{C}) \times (34.2 - 15.4)°\text{C} = \)

7. Calculate the mass of each reactant and product from the chemical amount shown in the following equations, and show how your calculations agree with the law of conservation of mass:
   (a) \( \text{H}_2(g) + \text{C}_6\text{H}_5\text{Cl}(g) \rightarrow 2 \text{ HCl}(g) \)
   (b) \( 2 \text{ CH}_3\text{OH}(l) + 3 \text{ O}_2(g) \rightarrow 2 \text{ CO}_2(g) + 4 \text{ H}_2\text{O}(g) \)
2.5 Classifying Chemical Reactions

By analyzing the evidence obtained from many chemical reactions, it is possible to distinguish patterns. On the basis of these patterns, certain generalizations about reactions can be formulated. The generalizations in Table 1 are based on extensive evidence and provide an empirical classification of most, but not all, common chemical reactions. The five types of reactions are described in the sections that follow. (For now, any reactions that do not fit these categories are classified as “other.”)

<table>
<thead>
<tr>
<th>Reaction type</th>
<th>Generalization</th>
</tr>
</thead>
<tbody>
<tr>
<td>formation</td>
<td>elements → compound</td>
</tr>
<tr>
<td>simple decomposition</td>
<td>compound → elements</td>
</tr>
<tr>
<td>complete combustion</td>
<td>substance + oxygen → most common oxides</td>
</tr>
<tr>
<td>single replacement</td>
<td>element + compound → element + compound</td>
</tr>
<tr>
<td></td>
<td>(metal + compound → metal + compound)</td>
</tr>
<tr>
<td></td>
<td>(or nonmetal + compound → nonmetal + compound)</td>
</tr>
<tr>
<td>double replacement</td>
<td>compound + compound → compound + compound</td>
</tr>
</tbody>
</table>

Formation Reactions

A formation reaction is the reaction of two or more elements to form either an ionic compound (from a metal and a nonmetal) or a molecular compound (from two or more nonmetals). An example of a reaction forming an ionic compound is the reaction of magnesium and oxygen shown in Figure 1.

word equation: magnesium + oxygen → magnesium oxide

chemical equation: $2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2\text{MgO}(s)$

For chemical reactions producing molecular substances, the only products that you will be able to predict at this time are those whose formulas you memorized from Table 3 in Section 1.6 (page 34); for example, $\text{H}_2\text{O}$.

Simple Decomposition Reactions

A simple decomposition reaction is the breakdown of a compound into its component elements, that is, the reverse of a formation reaction. Simple decomposition reactions are important historically since they were used to determine chemical formulas. They remain important today in the industrial production of some elements from compounds available in the natural environment. A well-known example that is easy to demonstrate is the simple decomposition of water (Figure 2).

word equation: water → hydrogen + oxygen

chemical equation: $2\text{H}_2\text{O}(l) \rightarrow 2\text{H}_2(g) + \text{O}_2(g)$

Combustion Reactions

A complete combustion reaction is the burning of a substance with sufficient oxygen available to produce the most common oxides of the elements making up the substance that is burned. Some combustions, like those in a burning candle or an untuned
automobile engine, are incomplete, and also produce the less common oxides such as carbon monoxide. Combustion reactions (Figure 3) are exothermic; these reactions provide the major source of energy for technological use in our society.

To successfully predict the products of a complete combustion reaction, you must know the composition of the most common oxides. If the substance being burned contains

- carbon, then CO₂(g) is produced
- hydrogen, then H₂O(g) is produced
- sulfur, then SO₂(g) is produced
- nitrogen, then assume NO₂(g) is produced
- a metal, then the oxide of the metal with the most common ion charge is produced (Figure 3)

A typical example of a complete combustion reaction is the burning of butane, C₄H₁₀(g):

word equation: butane + oxygen → carbon dioxide + water
chemical equation: 2 C₄H₁₀(g) + 13 O₂(g) → 8 CO₂(g) + 10 H₂O(g)

You will learn more about combustion reactions in Section 9.6.

---

**Section 2.5 Questions**

1. Rewrite each of the following reactions as a word or balanced chemical equation, and classify each reaction as formation, simple decomposition, or complete combustion. Assume the SATP states of matter unless otherwise indicated. For example,

\[ 2 \text{Na(s)} + \text{Cl}_2(g) \rightarrow 2 \text{NaCl(s)} \text{ (formation)} \]

- sodium + chlorine → sodium chloride

(a) lithium oxide → lithium + oxygen
(b) 2 KBr(s) → 2 K(s) + Br₂(l)
(c) 6 K(s) + N₂(g) → 2 K₃N(s)
(d) magnesium oxide → magnesium + oxygen
(e) 16 Al(s) + 3 S₈(s) → 8 Al₂S₃(s)
(f) methane + oxygen → carbon dioxide + water (vapour)

2. For each of the following reactions,

- classify the reaction type as formation or simple decomposition,
- predict the product(s) of the reaction,
- complete and balance the chemical equation
- complete the word equation.

Assume the most common ion charges and that the products are at SATP.

(a) Since the Bronze Age (about 3000 B.C.E.), copper has been produced by heating the ore that contains CuO(s).

- copper(II) oxide →

(b) When aluminium reacts with air, a tough protective coating forms. This coating helps prevent acidic substances, such as soft drinks (Figure 4), from reacting with the acids and thereby corroding the aluminium.

\[ \text{Al(s)} + \text{O}_2(g) \rightarrow \]

(c) Sodium hydroxide can be decomposed into its elements by melting it and passing electricity through it.

\[ \text{NaOH}(l) \rightarrow \]

(d) Very reactive sodium metal reacts with the poisonous gas chlorine to produce an inert, edible chemical.

\[ \text{Na(s)} + \text{Cl}_2(g) \rightarrow \]

(e) A frequent technological problem associated with the operation of swimming pools is that copper pipes react with aqueous chlorine.

\[ \text{Cu(s)} + \text{Cl}_2(aq) \rightarrow \]
(f) A major scientific breakthrough occurred in 1807 when Sir Humphry Davy isolated potassium by passing electricity through molten (melted) potassium oxide.

$$K_2O(l) \rightarrow$$

(g) When zinc is exposed to oxygen, a protective coating forms on the surface of the metal. This reaction makes zinc coating of metals (galvanizing) a desirable process for resisting corrosion.

zinc + oxygen →

(h) Translate the last equation above into an English sentence. Include the chemical amounts in moles.

3. State the names and chemical formulas for the most common oxides of carbon, hydrogen, sulfur, nitrogen, and iron.

4. For each of the following complete combustion reactions, complete and balance the chemical equation or complete the word equation. Assume the pure state of matter at SATP unless otherwise indicated.

(a) In Canada, many homes are heated by the combustion of natural gas (assume methane).

$$CH_4(g) + O_2(g) \rightarrow$$

(b) Nitromethane, CH₃NO₂(l), is a fuel commonly burned in drag-racing vehicles.

nitromethane + oxygen →

(c) Mercaptans (assume C₄H₉SH(g)) are added to natural gas to give it a distinct odour. The mercaptan burns with the natural gas.

$$C_4H_9SH(g) + O_2(g) \rightarrow$$

(d) Ethanol from grain can be added to gasoline as a fuel and antifreeze. It burns along with the gasoline.

ethanol + oxygen →

(e) Write a balanced equation for (d), and then translate the equation into an English sentence. Include the chemical amounts and states of matter.

(f) Most automobiles currently burn gasoline (assume octane) as a fuel.

octane + oxygen →

5. Rewrite each of the following reactions as a word equation or a balanced chemical equation, and classify each reaction as formation, simple decomposition, or complete combustion. (Some reactions may have two classifications.)

(a) Electricity is used to produce elements from molten potassium bromide at a high temperature.

$$2 KBr(l) \rightarrow 2 K(l) + Br_2(g)$$

(b) Coal burns in a power plant to produce heat for generating electrical energy.

carbon + oxygen → carbon dioxide

(c) Gasoline antifreeze burns in an automobile engine.

methanol + oxygen → carbon dioxide + water

(d) Poisonous hydrogen sulfide from natural gas is eventually converted to elemental sulfur using this reaction as a first step in about 50 gas plants in Alberta.

$$2 H_2S(g) + 3 O_2(g) \rightarrow 2 SO_2(g) + 2 H_2O(g)$$

(e) Hydrogen gas may be the automobile fuel of the future.

hydrogen + oxygen → water

(f) Toxic hydrogen cyanide gas can be destroyed in a waste treatment plant, such as the one at Swan Hills, Alberta.

Four moles of hydrogen cyanide gas react with nine moles of oxygen gas to produce four moles of carbon dioxide gas, two moles of water vapour, and four moles of nitrogen dioxide gas.

6. Classify the following reactions as formation, simple decomposition, or complete combustion. Predict the products of the reactions, write the formulas and states of matter, and balance the reaction equations:

(a) Al(s) + F₂(g) →

(b) NaCl(s) →

(c) S₈(s) + O₂(g) →

(d) methane + oxygen →

(e) aluminium oxide →

(f) propane burns

(g) C₄H₁₀(g) + O₂(g) →

7. Describe a technological application for two of the chemical reactions in question 6.

8. Write a brief empirical description of reactants and products for two of the chemical reactions in question 6.

9. List two benefits and two risks of using combustion reactions. Use examples of a fuel used in your area. Try to use a variety of perspectives in your answer (see Section 2.1, page 45).

Extension

10. Hydrogen-burning cars may become common in the future. Write perspective statements, pro and/or con, to the resolution that most cars should be burning hydrogen in twenty years. Provide at least one statement from each of scientific, technological, ecological, economic, and political perspectives.

[www.science.nelson.com](http://www.science.nelson.com)

11. Dr. John Polanyi is a Canadian Nobel laureate. One reaction that he studied was that of atomic hydrogen gas, H(g), with chlorine gas. Read the story on the Web site and write the chemical equation for this reaction.

[www.science.nelson.com](http://www.science.nelson.com)
2.6

Chemical Reactions in Solution

The reactions reviewed so far involve pure substances. The remaining two reaction types, single and double replacements, usually occur in aqueous solutions. As you know, substances dissolved in water are indicated by (aq).

A solution is a homogeneous mixture (page 12) of a solute (the substance dissolved) and a solvent (the substance, usually a liquid, that does the dissolving). Figure 1 shows a common example involving table salt and water. The solubility of a substance, which is covered in more detail in Chapter 5, is the maximum quantity of the substance that will dissolve in a solvent at a given temperature. For substances like sodium chloride (in table salt), the maximum quantity that dissolves in certain solvents is large compared with other solutes. Such solutes are said to be very soluble. When very soluble substances are formed as products in a single or double replacement reaction, the maximum quantity of solute that can dissolve is rarely reached; thus, the new substance remains in solution, and an (aq) notation is appropriate. Other substances, such as calcium carbonate (in limestone and chalk), are only slightly soluble. When these substances are formed in a chemical reaction, the maximum quantity that can dissolve is usually reached and most of this substance settles to the bottom as a solid. Solid substances formed from reactions in solution are known as precipitates (Figure 2, on the next page). They are indicated in a chemical reaction equation by (s).

A solubility chart outlines solubility generalizations for a large number of ionic compounds; see Table 1. A major purpose of this chart is to predict the state of matter for ionic compounds formed as products of chemical reactions in solution. This summary of solubility evidence is listed in two categories—very soluble (aq) (for example, sodium chloride) and slightly soluble (s) (for example, calcium carbonate). The solubility of ionic compounds in water can be predicted from the solubility chart.

At this point, you will not be expected to predict the solubility of molecular compounds in water, but you should memorize the examples in Table 2. Some elements, like the alkali metals, react with water, but most elements do not react or dissolve in water to any noticeable extent. In general, if an element is a reactant or product in a chemical reaction in an aqueous solution then assume its pure state of matter unless otherwise indicated.

---

**Table 1 Solubility of Ionic Compounds at SATP—Generalizations**

<table>
<thead>
<tr>
<th>Ion</th>
<th>Cl(^{−})</th>
<th>Br(^{−})</th>
<th>I(^{−})</th>
<th>S(^{2−})</th>
<th>OH(^{−})</th>
<th>SO(_{4}^{2−})</th>
<th>CO(_{3}^{2−})</th>
<th>PO(_{4}^{3−})</th>
<th>SO(_{3}^{2−})</th>
<th>CH(_{3})COO(^{−})</th>
<th>NO(_{3}^{−})</th>
<th>ClO(_{4}^{−})</th>
<th>H(<em>{2})O(^{+}) (H(</em>{3})O(^{+}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>very soluble (aq)</td>
<td></td>
<td></td>
<td></td>
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<tr>
<td>≥ 0.1 mol/L</td>
<td>most</td>
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</tr>
<tr>
<td>Group 1, NH(_{4})^{+}</td>
<td>Group 1, NH(_{4})^{+}</td>
<td>Ba(^{2+})</td>
<td>TI(^{+})</td>
<td></td>
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<td>slightly soluble (s)</td>
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</tr>
<tr>
<td>&lt; 0.1 mol/L (at SATP)</td>
<td>Ag(^{+}), Pb(^{2+}), Ti(^{+}), Hg(_{2}^{2+}), Cu(^{+})</td>
<td>most</td>
<td>most</td>
<td></td>
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<td></td>
</tr>
<tr>
<td>Ag(^{+}), Pb(^{2+}), Ca(^{2+}), Ba(^{2+}), Sr(^{2+}), Ra(^{2+})</td>
<td>most</td>
<td>Ag(^{+}), Pb(^{2+}), Ca(^{2+}), Ba(^{2+}), Sr(^{2+}), Ra(^{2+})</td>
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<tr>
<td>NH(_{4})^{+}</td>
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</tr>
<tr>
<td>Group 1, NH(_{4})^{+}</td>
<td>Group 1, NH(_{4})^{+}</td>
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<td>all</td>
<td>all</td>
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</tbody>
</table>

*Although these are particularly reliable, all generalizations have exceptions. This textbook specifically identifies any reference to an ionic compound solubility that is an exception to these generalizations.*
Single Replacement Reactions

A single replacement reaction is the reaction of an element with a compound to produce a new element and an ionic compound. This reaction usually occurs in aqueous solutions. For example, silver can be produced from copper and a solution of silver ions:

\[
\text{Cu(s)} + 2\text{AgNO}_3(\text{aq}) \rightarrow 2\text{Ag(s)} + \text{Cu(NO}_3)_2(\text{aq})
\]

\(\text{copper} + \text{silver nitrate} \rightarrow \text{silver} + \text{copper(II) nitrate}\)

Iodine can be produced from chlorine and aqueous sodium iodide:

\[
\text{Cl}_2(\text{g}) + 2\text{NaI(aq)} \rightarrow \text{I}_2(\text{s}) + 2\text{NaCl(aq)}
\]

\(\text{chlorine} + \text{sodium iodide} \rightarrow \text{iodine} + \text{sodium chloride}\)

We can predict the very soluble (aq) states for the two ionic products, copper(II) nitrate and sodium chloride, from the solubility chart (Table 1, page 61). Evidence shows that a metal replaces a metal ion to liberate a different metal as a product (as in the first preceding example) and a nonmetal replaces a nonmetal ion to liberate a different nonmetal as a product (as in the second example). Reactive metals, such as those in Groups 1 and 2, react with water to replace the hydrogen, forming hydrogen gas and a hydroxide compound. (In these reactions, hydrogen acts like a metal.)

Double Replacement Reactions

A double replacement reaction can occur between two ionic compounds in solution. In the reaction, the ions, by analogy, “change partners” to form the products. If one of the products is slightly soluble, it may form a precipitate, as shown in Figure 4. As the
term implies, **precipitation** is a double replacement reaction in which a precipitate forms. For example:

\[
\text{CaCl}_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + 2\text{NaCl}(\text{aq})
\]

compound + compound → compound + compound

Remember that chemists have created solubility generalizations to help us organize our knowledge. The generalizations serve our purpose and are only approximations of what is found in nature. There are always exceptions to generalizations.

In another kind of double replacement reaction, an acid reacts with a base, producing water and an ionic compound. This kind of double replacement reaction is known as **neutralization**. The reaction between hydrochloric acid and potassium hydroxide is an example:

\[
\text{HCl}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{KCl}(\text{aq})
\]

acid + base → water + ionic compound (a salt)

When writing chemical equations for both precipitation and neutralization reactions, consult the solubility chart (**Table 1**, page 61) to determine the state of matter of the ionic products. For neutralization reactions, it may be easier to balance the equation if you temporarily write the chemical formula for water as HOH(l) rather than H₂O(l).

**SUMMARY**

**Predicting Chemical Reactions**

<table>
<thead>
<tr>
<th>Type of reaction</th>
<th>Generalization</th>
<th>Notes</th>
</tr>
</thead>
<tbody>
<tr>
<td>formation</td>
<td>element + element → compound</td>
<td>metal + nonmetal → ionic compound</td>
</tr>
<tr>
<td></td>
<td></td>
<td>nonmetal + nonmetal → molecular compound</td>
</tr>
<tr>
<td>simple decomposition</td>
<td>compound → element + element</td>
<td></td>
</tr>
<tr>
<td>complete combustion</td>
<td>element or compound + oxygen → oxide</td>
<td></td>
</tr>
<tr>
<td>single replacement</td>
<td>element + compound → element + compound</td>
<td>metal + compound → metal + compound</td>
</tr>
<tr>
<td></td>
<td></td>
<td>nonmetal + compound → nonmetal + compound</td>
</tr>
<tr>
<td></td>
<td></td>
<td>metal + acid → hydrogen + compound</td>
</tr>
<tr>
<td>double replacement</td>
<td>compound + compound → compound + compound</td>
<td>precipitation reaction:</td>
</tr>
<tr>
<td></td>
<td></td>
<td>solution + solution → precipitate + solution</td>
</tr>
<tr>
<td></td>
<td></td>
<td>neutralization reaction:</td>
</tr>
<tr>
<td></td>
<td></td>
<td>acid + base → water + aqueous ionic compound</td>
</tr>
</tbody>
</table>

To write the correct chemical equation for a reaction, follow these steps:

Step 1: Use the reaction generalizations to classify the reaction.

Step 2: Use the reaction generalizations to predict the products of the chemical reaction and write the chemical equation.

(a) Predict, from theory, the chemical formulas for ionic compounds, and write the formulas from memory for molecular compounds and elements.

(b) Include states of matter, using the rules and generalizations.

Step 3: Balance the equation without changing the chemical formulas.

**EXTENSION**

**Hard and Soft Water**

Precipitation reactions affect our water supply. Extend your understanding of precipitation by hearing about the everyday chemistry of hard and soft water.

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Section 2.6 Questions

1. The following chemical reactions occur in a water environment. Write the balanced chemical equation, including the states of matter, (s) or (aq), for each reaction:
   (a) lead(II) nitrate + lithium chloride → lead(II) chloride + lithium nitrate
   (b) ammonium iodide + silver nitrate → silver iodide + ammonium nitrate
   (c) The net reaction during the discharge cycle of a car battery is one mole of lead and one mole of solid lead(IV) oxide reacting with two moles of sulfuric acid to produce two moles of water and two moles of lead(II) sulfate.

2. The following reactions occur in a water environment. Write the balanced chemical equation, including SATP states of matter, for each chemical.
   (a) Copper can be extracted from solution by reusing cans that contain iron.
      iron + copper(II) sulfate → copper + iron(III) sulfate
   (b) Water can be clarified by producing a gelatinous precipitate.
      aluminium sulfate + calcium hydroxide → aluminium hydroxide + calcium sulfate
   (c) Chlorine is used to extract bromine from sea water.
      Cl₂(g) + NaBr(aq) → Br₂(l) + NaCl(aq)
   (d) During photosynthesis in a plant, carbon dioxide reacts with water to produce glucose and oxygen.
      carbon dioxide + water → glucose + oxygen

3. Use the solubility table (Table 1, page 61), the generalization that all elements (except chlorine) are slightly soluble in water, and the molecular solubility in Table 2 to predict the solubility of the following chemicals in water. Classify the chemical and then write the chemical formula with (aq) to indicate that the chemical is very soluble and with the pure state of matter, (s), (l), or (g), to indicate that the chemical is slightly soluble.
   (a) Zn (dry cell container and reactant)
   (b) P₄ (white phosphorus)
   (c) Ca₃(PO₄)₂ (sugar)
   (d) methanol (windshield and gasoline antifreeze)
   (e) octane (gasoline component)
   (f) barium sulfate (gastric X-rays)
   (g) sodium hydroxide (drain cleaner)
   (h) ammonia (window and general cleaner)
   (i) hydrogen fluoride (used to etch glass)

4. Communication systems are very important in chemistry. Describe the difference in what is being communicated in the following sets of symbols.
   (a) P₄ and 4 P
   (b) Cl₂(g) and Cl(g)
   (c) H₂O(l) and H₂O(g)
   (d) H₂O₂ and H₂ + O₂
   (e) NaCl(s) and NaCl(aq)
   (f) Mg and Mg²⁺

5. For each of the following reactions, classify the reaction, predict the products of the reaction, and complete and balance the chemical equation. (Assume the most common ion charge and state at SATP if not indicated otherwise.)
   (a) Expensive silver metal is recovered in a lab by placing inexpensive aluminium foil in aqueous silver nitrate.
      Al(s) + AgNO₃(aq) →
   (b) When aqueous potassium hydroxide is added to a well-water sample, the formation of a rusty-brown precipitate indicates the presence of an iron(III) compound in the water.
      KOH(aq) + FeCl₃(aq) →
   (c) A chemist in a consumer-protection laboratory adds aqueous sodium hydroxide to determine the concentration of acetic acid, CH₃COOH(aq), in a vinegar sample.
      CH₃COOH(aq) + NaOH(aq) →
   (d) A dishonest 16th-century alchemist, who tried to fool people into believing that iron could be changed into gold, dipped an iron strip into aqueous copper(II) sulfate.
      Fe(s) + CuSO₄(aq) →
   (e) Translate equation (d) into an English sentence. Include the chemical amounts and states of matter.

6. Complete the Prediction and diagnostic tests of the investigation report. Write up the diagnostic tests, including any controls, as part of the Design.

   Problem
   What are the products of the reaction of sodium metal and water?

   Design
   A very small piece of sodium metal is placed in distilled water and some diagnostic tests are carried out to identify the products.

7. State diagnostic tests for each of the product(s) in the following chemical reactions. Use the “If (procedure) and (evidence), then (analysis)” format.
   (a) 2 H₂O(l) → 2 H₂(g) + O₂(g)
   (b) H₂(g) + Cl₂(aq) → 2 HCl(aq)
   (c) NH₄(g) + H₂O(l) → NH₃OH(aq)

Extension

8. Critiquing and creating experimental designs are important skills for scientific literacy. Create experimental designs to test at least two of the claims made by the following individuals.
   (a) A salesperson claims that wrapping magnets around water pipes will reduce the amount of hard-water scaling that accumulates on the inside of the pipes.
   (b) A psychic claims that he can see halos over the heads of some identified individuals in the audience and not over others.
   (c) A psychic claims to be able to bend spoons with his mind—a feat of psychokinesis.
   (d) A salesperson claims that a copper bracelet relieves pain in the wrist.
Outcomes

Knowledge

• use kinetic molecular theory and collision theory to explain how chemical reactions occur (2.2)
• write balanced chemical equations (2.2, 2.3)
• interpret balanced chemical equations in terms of chemical amount (in moles) (2.3)
• convert between chemical amount and mass (2.4)
• classify chemical reactions (2.5, 2.6)
• predict the solubility of elements and ionic and molecular compounds in water (2.6)
• predict products for chemical reactions (2.5, 2.6)

STS

• state the technological application of important chemicals and chemical reactions (2.1, 2.3, 2.4, 2.5, 2.6)
• identify risks and benefits of some important chemical reactions (2.1, 2.3, 2.5)

Skills

• read and write laboratory reports (2.6)
• create and critique experimental designs (2.6)

Key Terms

2.1
STS
perspective
scientific
technological
ecological
economic
political

2.2
physical change
chemical change
nuclear change
kinetic molecular theory
diagnostic test
law of conservation of mass
balanced chemical equation
coefficient

2.3
amount of substance
chemical amount
Avogadro’s number
mole

2.4
molar mass

2.5
formation reaction
simple decomposition reaction
complete combustion reaction

2.6
solute
solvent
solubility
precipitate
single replacement reaction
double replacement reaction
precipitation
neutralization

MAKE a summary

1. Use the Key Terms to prepare concept maps that are centred on chemical reactions. Include, along with an example of each:
   (a) evidence for the occurrence of chemical reactions
   (b) classes of chemical reactions
   (c) solubility of elements, ionic compounds, and molecular compounds
   (d) perspectives on an STS issue
2. Search for ways to link the concept maps prepared in the previous question.
3. Refer back to your answers to the Starting Points questions at the beginning of this chapter. How has your thinking changed?

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Science and the Courts

Scientific evidence is increasingly being introduced into legal cases: DNA, blood alcohol levels, psychological profiles—all are intended to support either the prosecution’s or the defence’s case. But are the “facts” beyond doubt? Three experts give their opinions about interpreting scientific evidence for the courts.

www.science.nelson.com
Many of these questions are in the style of the Diploma Exam. You will find guidance for writing Diploma Exams in Appendix H. Exam study tips and test-taking suggestions are on the Nelson Web site. Science Directing Words used in Diploma Exams are in bold type.

DO NOT WRITE IN THIS TEXTBOOK.

Part 1

1. Pentane, $C_5H_{12}(g)$, is a major component of naptha, the fuel manufactured by chemical engineers for camp stoves (Figure 1). The complete combustion of pentane is represented by the chemical equation

$$C_5H_{12}(g) + xO_2(g) \rightarrow yCO_2(g) + zH_2O(g)$$

When the equation is balanced, the coefficients are $x$, $y$, $z$.

![Figure 1](Naphtha is a popular camp fuel because it is easy to transport. It does, however, require careful handling, as the HHP symbols indicate.)

2. Environmental chemists have found that nitrogen oxides, $NO_x$, can cause acid rain, as shown by the chemical equation

$$NO_x(g) + yH_2O(l) \rightarrow zHNO_3(aq) + pNO(g)$$

When the equation is balanced the coefficients are $x$, $y$, $z$, $p$.

3. Classification systems help us organize our knowledge. A reaction in which a precipitate is formed when two solutions of ionic compounds are mixed is classified as a

A. formation reaction
B. single replacement reaction
C. double replacement reaction
D. simple decomposition reaction

4. Research indicates that manufactured chemicals are among the most persistent pollutants on Earth. Scientists have found some chemicals in the tissues of polar bears, seals, tropical birds, dolphins, and humans. The primary perspective of this research is

A. technological
B. ecological
C. economic
D. scientific

5. The concept that the smallest entities of a substance are in continuous motion is central to

A. Dalton’s atomic theory
B. Rutherford’s atomic theory
C. the kinetic molecular theory
D. the law of conservation of mass

6. Chemists classify an acid–base neutralization reaction as a

A. formation reaction
B. single replacement reaction
C. double replacement reaction
D. simple decomposition reaction

Part 2

7. Use the solubility table (Table 1 on page 61), the generalization that all elements (except chlorine) are only slightly soluble in water, and the molecular solubilities in Table 2 to predict the solubility of the following chemicals in water. Classify the chemical and then write the chemical formula with (aq) if chemical is very soluble and with the pure state of matter, (s), (l), or (g), if the chemical is slightly soluble in a water environment.

(a) sucrose (table sugar)
(b) methane (natural gas)
(c) calcium sulfate (gypsum)
(d) carbon (charcoal/graphite)
(e) sulfuric acid (car batteries)
(f) sodium carbonate (water softener)
(g) ammonium nitrate (fertilizer)
(h) sulfur (from H$_2$S in sour gas)
(i) silver bromide (photographic film)
(j) magnesium hydroxide (milk of magnesia)

8. For each of the following word equations, write a balanced chemical equation, and classify the reaction.

(a) aqueous sodium hydroxide + sulfuric acid → water + sodium sulfate
(b) propane + oxygen → carbon dioxide + water
(c) aluminium + aqueous copper(II) chloride → copper + aluminium chloride
(d) molten sodium hydroxide → sodium + oxygen + hydrogen
(e) calcium + chlorine → calcium chloride
(f) aqueous lead(II) nitrate + aqueous sodium chloride → lead(II) chloride + sodium nitrate

9. Classify each of the following reactions as formation, simple decomposition, complete combustion, single replacement, or double replacement. Predict the products of the reactions, write the chemical formulas and states of matter, and balance the reaction equations.

(a) $KCl(s) \rightarrow$
(b) $Cu(s) + Cl_2(g) \rightarrow$
(c) $C_6H_{12}(g) + O_2(g) \rightarrow$
(d) $AgNO_3(aq) + NaCl(aq) \rightarrow$
(e) $Al(s) + Cu(NO_3)_2(aq) \rightarrow$
(f) $C_3H_8(l) + O_2(g) \rightarrow$
(g) $Al_2O_3(s) \rightarrow$
(h) Fe(s) + Br₂(l) →
(i) Cu(NO₃)₂(aq) + NaOH(aq) →
(j) H₃PO₄(aq) + Ca(OH)₂(aq) →

10. Translate each of the following balanced equations into an English sentence. Include the chemical amounts and the states of matter for all the substances involved.
   (a) 4 NH₃(g) + 7 O₂(g) → 4 NO₂(g) + 6 H₂O(g)
   (b) 3 CaCl₂(aq) + 2 Na₃PO₄(aq) → Ca₃(PO₄)₂(s) + 6 NaCl(aq)
   (c) 2 NaCl(l) → 2 Na(l) + Cl₂(g)

11. Describe diagnostic tests for each of the product(s) in the following chemical reactions. Use the "If (procedure) and (evidence), then (analysis)" format.
   (a) 2 K(s) + 2 H₂O(l) → 2 KOH(aq) + H₂(g)
   (b) SO₃(g) + H₂O(l) → H₂SO₄(aq)
   (c) Cl₂(g) + 2 NaI(aq) → 2 NaCl(aq) + I₂(s)

12. Complete the Prediction and diagnostic tests of the investigation report. Describe the diagnostic tests, including any controls, as part of the Design. Use the "If (procedure) and (evidence), then (analysis)" format for the diagnostic tests.

**Problem**
What are the products of the reaction of iron metal and hydrochloric acid?

**Design**
A short piece of iron wire is placed in a dilute solution of hydrochloric acid and some diagnostic tests are carried out to identify the products.
Many of these questions are in the style of the Diploma Exam. You will find guidance for writing Diploma Exams in Appendix H. Exam study tips and test-taking suggestions are on the Nelson Web site. Science Directing Words used in Diploma Exams are in bold type.

DO NOT WRITE IN THIS TEXTBOOK.

Part 1
The modern periodic table was developed from evidence of periodicity in chemical and physical properties. The periodic table is an efficient means of organizing a vast body of empirical knowledge of the elements.

1. The scientist who created the first periodic table was
A. Niels Bohr  C. Albert Einstein
B. John Dalton  D. Dmitri Mendeleev

2. The family of elements listed in Group 2 of the periodic table is known as the
A. actinoids  C. alkaline-earth metals
B. alkali metals  D. transition elements

3. All the members of a certain family of elements are soft, silver-coloured conductors of electricity that react violently with water and form ions with a 1+ charge. The family name for this group of elements is
A. halogens  C. alkali metals
B. noble gases  D. alkaline-earth metals

4. All the members of a certain family of elements are very reactive and form ions with a 1– charge. The family name for this group of elements is
A. actinides  C. alkali metals
B. halogens  D. lanthanoids

5. The number of protons, electrons, and neutrons, respectively in an ion of lithium-7 are

6. The number of protons, electrons, and neutrons, respectively in an atom of carbon-14 are

7. Chemists often classify changes in matter as physical change, chemical change, or nuclear change. An example of a physical change is
A. formation  C. evaporation
B. combustion  D. neutralization

8. An equation that represents a physical change is
A. H₂O(l) → H₂O(l)
B. 2 H₂(g) + O₂(g) → 2 H₂O(g)
C. 2 H₂(g) + O₂(g) → 2 H₂O(l)
D. 2 H₂O(l) → 2 H₂(g) + O₂(g)

Use a periodic table to answer questions 2 to 6.

Use this information to answer questions 9 and 10. Some descriptors may be used more than once.

Physical and chemical changes can be described empirically. The empirical descriptions include:
1. odour change
2. state change
3. colour change

List, in numerical order, the description(s) that applies/apply to each of the following reactions.

9. 2 C₈H₁₈(l) + 25 O₂(g) → 16 CO₂(g) + 18 H₂O(g)

10. Zn(s) + CuSO₄(aq) → Cu(s) + ZnSO₄(aq)

11. Knowledge can be classified as
1. empirical
2. theoretical

Use the above classes of knowledge (i.e., 1 and 2) to classify the following statements, in order:
• A yellow precipitate formed.
• Atoms/ions were rearranged.
• Chemical bonds were broken and formed.
• The solution changed from purple to colourless.

12. The multiple-step industrial process for producing sodium carbonate uses common salt (sodium chloride) and limestone (calcium carbonate) as raw material.

_ NaCl(s) + _ CaCO₃(s) → _ CaCl₂(s) + _ Na₂CO₃(s)

The coefficients for balancing the overall reaction equation are, in order, ___  ___  ___  ___

13. Propane is widely used as fuel for heating and cooking.

_ C₃H₈(g) + _ O₂(g) → _ CO₂(g) _ H₂O(g)

The coefficients for balancing the reaction equation for the complete combustion of propane are, in order, ___  ___  ___  ___

14. A spontaneous chemical reaction occurs when a crumpled piece of aluminium foil is dropped into a beaker of aqueous copper(II) sulfate.

_ Al(s) + _ CuSO₄(aq) → _ Cu(s) + _ Al₂(SO₄)₃(aq)

The coefficients for the balanced reaction equation are, in order, ___  ___  ___  ___

15. An early industrial process for producing sodium metal involved the electrolysis of molten sodium hydroxide.

_ NaOH(l) → _ Na(l) + _ H₂(g) + _ O₂(g)

The coefficients for the balanced reaction equation are, in order, ___  ___  ___  ___
16. The basic SI unit for chemical amount is the
   A. gram
   B. litre
   C. metre
   D. mole

17. The units for the molar mass of a chemical are
   A. g/mol
   B. g/L
   C. mol/L
   D. mol/g

18. An alcohol lamp uses 2.50 mol of methanol, CH₃OH(l), while providing emergency lighting. The mass of methanol burned is ______________ g.

19. A student obtains 150 g of ammonium sulfate, (NH₄)₂SO₄(s), to prepare a fertilizer solution. The chemical amount of ammonium sulfate obtained is ______________ mol.

20. A laboratory technician needs 7.50 mmol of potassium permanganate, KMnO₄(s), to prepare a solution for use in a chemical analysis. The mass of potassium permanganate required is ______________ g.

21. The lead(II) ions in a waste laboratory solution precipitate as lead(II) carbonate, PbCO₃(s). If 285 g of dry lead(II) carbonate is produced, the chemical amount of lead(II) ions precipitated is ______________ mol.

Part 2

22. The periodic table summarizes and organizes a wealth of empirical and theoretical knowledge about chemical elements. Define the following terms associated with the periodic table:
   (a) group
   (b) period
   (c) staircase line

23. Copy and complete the following table of atoms and ions.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Name</th>
<th># protons</th>
<th># electrons</th>
<th>Net charge</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>sulfide ion</td>
<td></td>
<td>35</td>
<td>36</td>
</tr>
<tr>
<td></td>
<td>Ca²⁺</td>
<td>23</td>
<td>23</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>26</td>
<td></td>
<td>3⁺</td>
</tr>
<tr>
<td></td>
<td></td>
<td>18</td>
<td>0</td>
<td></td>
</tr>
</tbody>
</table>

24. Many radioisotopes are produced artificially for use in medical diagnosis or therapy. Copy and complete the following table of radioisotopes assuming they are used as atoms.

<table>
<thead>
<tr>
<th>Name</th>
<th>Use</th>
<th># protons</th>
<th># electrons</th>
<th># neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>cobalt-60</td>
<td>cancer treatment</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>hyperthyroid treatment</td>
<td>53</td>
<td>78</td>
<td></td>
</tr>
<tr>
<td></td>
<td>reduces white cell count</td>
<td>15</td>
<td>17</td>
<td></td>
</tr>
<tr>
<td>strontium-85</td>
<td>bone scanning</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

25. For each of the following word equations, write a balanced chemical equation and classify the reaction.
   (a) Butane is a convenient fuel for camping (Figure 1).
       butane + oxygen → carbon dioxide + water vapour

   (b) Beautiful crystals form when copper objects are immersed in silver nitrate solutions.
       copper + silver nitrate → copper(II) nitrate + silver

   (c) Toxic cadmium ions can be removed from industrial effluent by sodium carbonate.
       cadmium nitrate + sodium carbonate →

   (d) Sir Humphry Davy discovered potassium by using electricity to decompose molten potassium hydroxide.
       potassium hydroxide →

   (e) Very pure or freshly cleaned aluminium forms a protective oxide coating when it is exposed to air.
       aluminium + oxygen → aluminium oxide
26. Balance the following reaction equations, and then write the equation in words, including the states of matter and chemical amounts.
   (a) Methanol burners keep food warm at a buffet.
   \[ CH_3OH(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g) \]
   (b) Phosphate ions can be removed from solution by adding calcium chloride.
   \[ Na_3PO_4(aq) + CaCl_2(aq) \rightarrow NaCl(aq) + Ca_3(PO_4)_2(s) \]

27. Complete and then balance the following chemical equations.
   (a) Magnesium burns with a brilliant white light.
   \[ Mg(s) + O_2(g) \rightarrow \]
   (b) The industrial production of sodium metal involves using electricity to decompose molten sodium chloride.
   \[ NaCl(l) \rightarrow \]
   (c) Aluminium sulfate and sodium hydroxide solutions react to form a gelatinous precipitate.
   \[ Al_2(SO_4)_3(aq) + NaOH(aq) \rightarrow \]
   (d) When its protective coating is removed, aluminium reacts vigorously with hydrochloric acid.
   \[ Al(s) + HCl(aq) \rightarrow \]
   (e) Candles are included in emergency kits because they can produce both heat and light.
   \[ C_{25}H_{52}(s) + O_2(g) \rightarrow \]

28. Complete the Prediction and Design of the investigation report. Include three diagnostic tests in the Design to determine whether the predicted reaction has taken place and the predicted products have formed.

**Problem**
What are the products of the reaction of aqueous copper(II) chloride and sodium hydroxide solution?

29. For each of the following reactions, translate the information into a balanced reaction equation. Then classify the main perspective—scientific, technological, ecological, economic, or political—suggested by the introductory statement.
   (a) Oxyacetylene torches are used to produce high temperatures for cutting and welding metals such as steel (Figure 2). This process involves burning acetylene, \( C_2H_2(g) \), in pure oxygen.
   (b) During chemical research conducted in 1808, Sir Humphry Davy produced magnesium metal by decomposing molten magnesium chloride using electricity.
   (c) An inexpensive application of single replacement reactions uses scrap iron to produce copper metal from waste copper(II) sulfate solutions.
   (d) The emission of sulfur dioxide into the atmosphere creates problems between different levels of government, both nationally and internationally. Sulfur dioxide is produced when zinc sulfide is roasted in a combustion-like reaction in a zinc smelter.
   (e) Burning leaded gasoline added toxic lead compounds to the environment, which damaged both plants and animals. Leded gasoline contained tetraethyl lead, \( Pb(C_2H_5)_4(l) \), which undergoes a complete combustion reaction in a car engine.

30. Classify the perspective being communicated by each of the following statements about global warming.
   (a) An invention is needed to remove carbon dioxide from gases emitted from oil refineries and power plants.
   (b) More research is needed to confirm or refute the causes of global warming.
   (c) The cost for stopping global warming is enormous.
   (d) Votes can be won or lost over global warming.
   (e) Profits will be reduced if greenhouse gas emissions are to be reduced.

31. Critiquing and creating experimental designs are important skills for scientific literacy. Create experimental designs to test at least two of the claims made by the following individuals. (See Appendix B.4.)
   (a) A believer in the power of magnets claims that sleeping with flexible and padded magnets in your pillow provides a more restful sleep.
   (b) A group of disbelievers claims that the photos and video from the 1969 Apollo 11 landing on the moon are fake because the shadows created by the Sun are not parallel in the video and photos.
   (c) A psychic claims to be able to reproduce simple drawings made by a person who comes up from the audience.
   (d) An alternative medical care provider claims to be able to cure a disease that standard medical practices cannot.
   (e) A commercial for a shampoo claims that your hair will have more body if you use this shampoo.
   (f) A commercial claims that a particular detergent removes grass stains from clothes better than other leading detergents.

*Figure 2*
The flame of an oxyacetylene torch is hot enough to melt most metals.
32. Complete the Prediction (including possible diagnostic tests), Analysis (including reaction types), and Evaluation (Parts 2 and 3) for the following investigation report.

**Purpose**
To test the single and double replacement reaction generalizations.

**Problem**
What reaction products are formed when the following substances are mixed?

(a) aqueous chlorine and potassium iodide solution
(b) solutions of magnesium chloride and sodium hydroxide
(c) solutions of aluminium nitrate and sodium phosphate
(d) magnesium metal and hydrochloric acid
(e) sodium hydroxide solution and chromium(III) chloride solution
(f) lithium metal and water
(g) a clean cobalt strip and a silver nitrate solution
(h) nitric acid and an ammonium acetate solution

**Design**
Diagnostic test information such as evidence of chemical reactions (Table 2), ion colours and solubilities (reference tables, inside back cover), and specific tests for products (Appendix C.3), are predicted, for convenience, along with the balanced chemical equations. The general plan is to observe the substances before and after mixing, and conduct the appropriate diagnostic tests.

**Evidence**

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Observations</th>
</tr>
</thead>
</table>
| (a)      | • The colourless solutions produced a yellow-brown colour when mixed (Figure 3).  
• A violet-purple colour appeared in the chlorinated layer when a hydrocarbon was added. |
| (b)      | • The colourless solutions produced a white precipitate when mixed. |
| (c)      | • The colourless solutions produced a white precipitate when mixed. |
| (d)      | • The silvery solid added to the colourless solution produced gas bubbles and a green solution.  
• The gas produced a pop sound when ignited. |
| (e)      | • The colourless sodium hydroxide and green chromium(III) chloride solutions produced a dark precipitate and a colourless solution. |
| (f)      | • The soft, silvery solid and colourless liquid produced gas bubbles and a colourless solution (Figure 4).  
• The gas produced a pop sound when ignited.  
• Red litmus turned blue in the final solution.  
• The final solution produced a bright red flame colour. |
| (g)      | • The silvery solid and colourless solution produced a pink solution and silvery needles. |
| (h)      | • The colourless solutions remained colourless when mixed.  
• A vinegar odour was produced. |

*Figure 3*  
Chlorine reacting with potassium iodide.

*Figure 4*  
The reaction of lithium with water.