unit

4 Quantitative Relationships in Chemical Changes

Chemical substances are part of our lives at every level—medicines, fuels, plastics, fertilizers to grow our food; the list seems endless. This unit focuses on how chemistry and technology interact to allow us to identify, measure, produce, and use chemical substances.

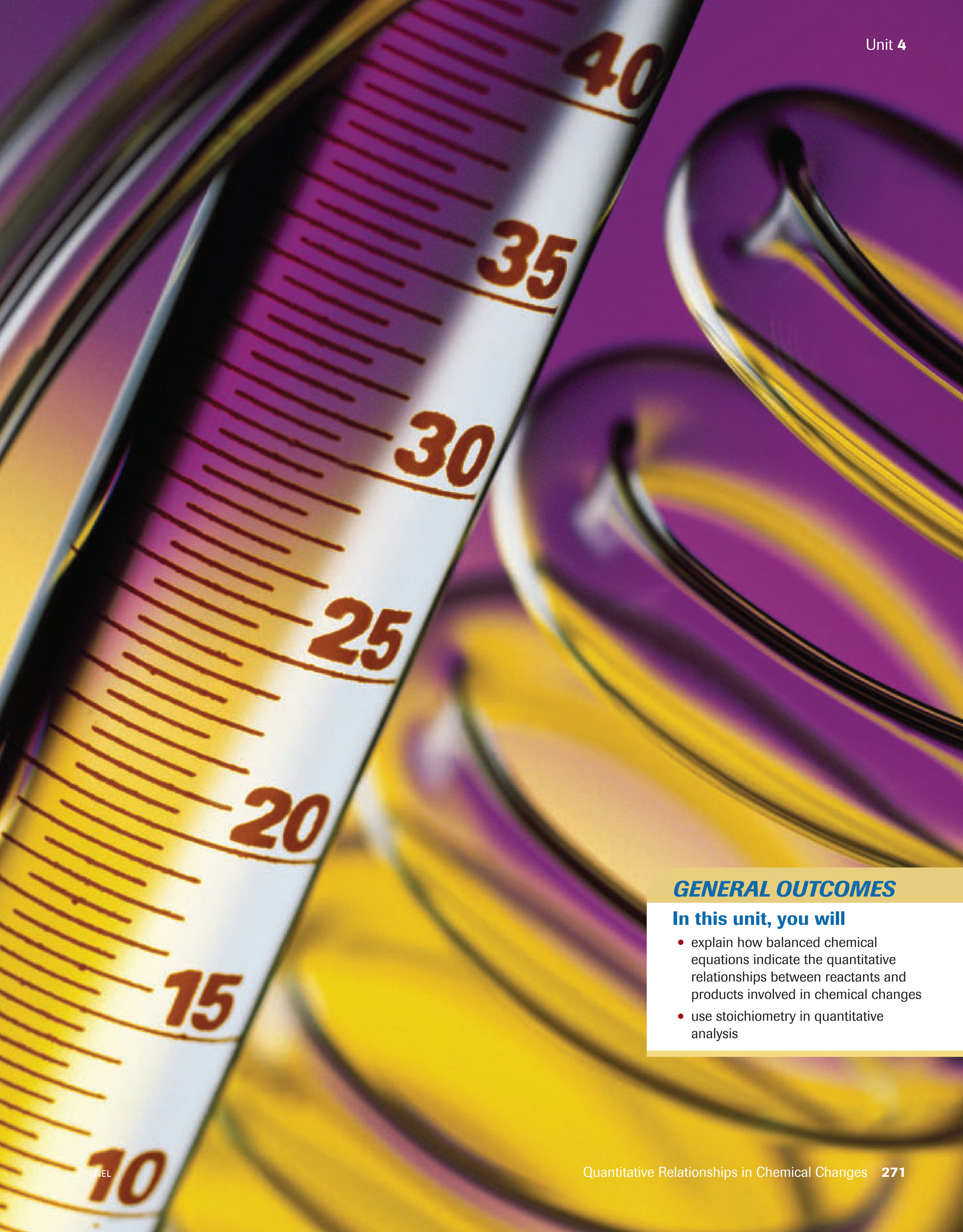
Quantitative relationships in chemical processes have always been important to humankind. One of the earliest written records we have (preserved on clay tablets) is a pharmacopeia, on which a Sumerian physician recorded the proper amounts of ingredients for making prescriptions—more than 4000 years ago! Interestingly, one ingredient listed often is the ionic compound, potassium nitrate. This shows that the Sumerians had some knowledge of basic chemistry at that time. Similarly, food recipes emphasize the critical importance of using proper quantities. Oral traditions ensured that Canada's Aboriginal peoples retained knowledge of proper proportions of ingredients for making the essential high-energy and long-lasting food staple called pemmican.

As science explored the quantities and proportions of substances that were involved in chemical reactions, technology developed to create, monitor, and control processes using them. This technology is applied by industry, commerce, and consumers to improve quality of life and to solve problems.

When science and technology develop, there are always more problems and questions created by the new information, processes, and skills. Understanding is the key to being able to evaluate the risks and benefits of chemical substances, and this begins with knowledge of how to identify and measure quantities of substances in chemical reactions.

As you progress through the unit, think about these focusing questions:

- How do scientists, engineers, and technologists use mathematics to analyze chemical changes?
- How are balanced chemical equations used to predict yields in chemical reactions?



GENERAL OUTCOMES

In this unit, you will

- explain how balanced chemical equations indicate the quantitative relationships between reactants and products involved in chemical changes
- use stoichiometry in quantitative analysis

Unit 4

Quantitative Relationships in Chemical Changes

ARE YOU READY?

Prerequisites

Concepts

- chemical formulas
- balanced reaction equations
- dissociation and ionization
- molar mass
- molar volume
- amount concentration
- ideal gas law
- acids and bases
- pH
- indicators

Skills

- WHMIS
- SI notation
- diagnostic tests

You can review prerequisite concepts and skills on the Nelson Web site, in the Chemistry Review unit, and in the Appendices.

A Unit Pre-Test is also available online.

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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. For each of the following combinations of reagents, predict the products, and write a complete, balanced chemical reaction equation:
   (a) Calcium chloride and sodium carbonate solutions are mixed (**Figure 1**).
   (b) Zinc metal reacts in hydrochloric acid (**Figure 2**).
   (c) Water undergoes simple decomposition.
   (d) Iron burns in pure oxygen (**Figure 3**).
   (e) Copper metal is placed in an aqueous solution of silver nitrate (**Figure 4**).

Figure 1

Figure 2

Figure 3

Figure 4

2. Many reactions will only occur in aqueous solution. Assume that each of the substances below is placed in water. Rewrite the formula for each substance, including the physical state, to indicate whether it has high or low solubility in water at SATP. Where appropriate, write a dissociation or ionization equation.
   (a) $Ca(NO_3)_2$
   (b) $PbCl_2$
   (c) HCl
   (d) NaOH

3. As part of an environmental analysis, the pH of a sample of lake water was measured to be 4.8.
   (a) Calculate the hydronium ion amount concentration in the lake water.
   (b) What would be the colours of the following indicators if placed in this lake water: methyl orange, bromothymol blue, phenolphthalein?

4. The molar mass of any substance is a useful conversion factor because it allows us to understand laboratory measurements of mass in terms of the chemical amount and vice versa. Copy and complete **Table 1**.

Table 1 Mass and Amount Conversions for Solids or Liquids

Formula	Mass (g)	Chemical amount (mol)
$(NH_4)_3PO_4(s)$	44.00	
$CH_3COOH(l)$		0.058

5. For reactions that occur in solution, the most useful conversion factor is the concentration, expressed as an amount concentration. Copy and complete **Table 2**.

Table 2 Volume, Concentration, and Amount Conversions for Solutions

Formula	Volume (L)	Amount concentration (mol/L)	Chemical amount (mol)
NaOH(aq)	2.20	0.500	
HCl(aq)		11.6	0.0400
$Na_2SO_4(aq)$	0.655		0.740

6. Calculating chemical amounts for gases often involves a number of variables. Sometimes a molar volume at STP or SATP can be used, but more often the ideal gas law is required. Copy and complete **Table 3**.

Table 3 Pressure, Temperature, Volume, and Amount Conversions for Gases

Formula	Pressure (kPa)	Temperature (K)	Volume (L)	Amount (mol)
$CH_4(g)$	101.325 (exact)	273.15 (exact)	13.7	
$UF_6(g)$	400	400	1.00	
$CO_2(g)$	100	298		2.0
Ar(g)	100	298	4.00	

STS Connections

7. Provide one example of a chemical product or process that shows how chemical technology is used to solve practical problems.

Skills

8. Write a brief experimental design, including diagnostic tests, to distinguish between neutral ionic, neutral molecular, acidic, and basic solutions.
9. Match each WHMIS symbol in **Table 4** to what it represents in the second column.
10. Identify the substance in each test shown in **Figure 5**.

Table 4 Matching WHMIS Symbols

Symbol	Class: Category
(a) 	Class B: Flammable and Combustible Materials
(b) 	Class C: Oxidizing Materials
(c) 	Class D: Toxic Materials Immediate and Severe
(d) 	Class F: Dangerously Reactive Materials

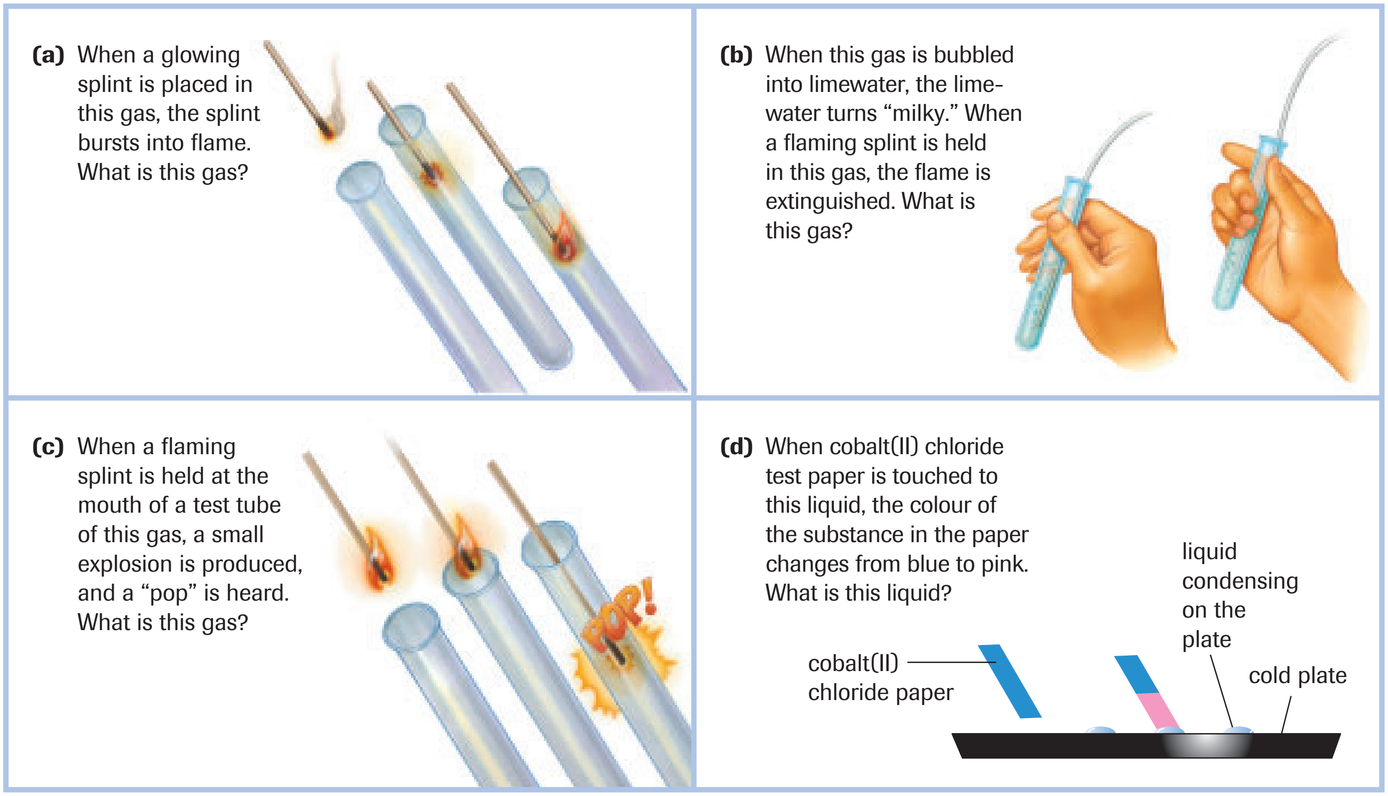

Figure 5
Diagnostic tests

chapter

7 Stoichiometry

In this chapter

Understanding quantitative relationships of reactants and products in chemical reactions is very important in chemical technology. The industrial production of fertilizer (**Figure 1**), the combustion of fuels, the treatment of water, and even our personal use of antacids are just a few examples. For each of these examples it is necessary to understand reacting quantities in order to understand the technology behind these products and processes.

In a general sense, food preparation is also chemical technology—products are produced from raw materials, using processes that are often based on chemical reactions. A recipe is a procedure that includes specific quantities and steps to be used to obtain the desired product. Characteristic of all technologies, you do not have to understand the science behind the process to be successful. However, if you want to explain the technology, the chemistry concepts become very important.

The laboratory study of chemical reactions requires simple technology: Chemists need to be able to identify the products when known substances react, often using the diagnostic tests that you already know. The study of chemical quantities used and/or produced in a reaction requires slightly more sophisticated technology and the skill to use it to make accurate measurements. A prediction of quantities also depends on scientific knowledge, such as balanced chemical reaction equations, chemical amounts, and their relationship to the chemical equation.

In this chapter, you will build on your previous knowledge of chemical formulas, chemical equations, and amount calculations. You will study reactions to interpret reaction equations and to calculate and predict the effects of controlling quantities of chemicals involved in a reaction.

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. Are all individual entities in the reactants important when a chemical reaction occurs?
2. What is the simple relationship between the quantities of any two reactants and products in a chemical reaction?
3. Compare the procedures used for measuring and calculating the quantities of solids, liquids, gases, and solutions involved in chemical reactions.

Career Connections:
Aerospace Engineer; Chemical Engineer; Chemical Technologist; Soil Scientist

Figure 1
This Agrium™ industrial plant site near Redwater, Alberta, combines science and technology to produce nitrogen-based fertilizers using natural gas. Agrium™ operates other Alberta plants at Carseland, Joffre, and Fort Saskatchewan. These plants produce a gross combined total of nearly two and a half million tonnes of ammonia per year, as well as many other highly soluble nitrogen-rich compounds.

Exploration — The Problem Is What You Don't See!

For many chemical reactions, the substances involved are invisible. This presents unique challenges in detecting, measuring, and calculating amounts of chemicals that you might not even notice. A balanced equation written for the combustion reaction of a typical hydrocarbon compound found in candle wax, $C_{17}H_{36}$, involves three substances that are invisible.

Materials: wax candle (about 2 cm in diameter), 250 mL Erlenmeyer flask, stopper to fit the flask, 600 mL beaker, cobalt(II) chloride test paper, limewater solution, knife

- Write and balance a reaction equation that describes the reaction of "burning" the candle.
- Cut the candle bottom level, so that when the candle is sitting on a flat surface, the candle flame height will be about 3 cm lower than the height of the 600 mL beaker.
- Light the candle. When it is burning well, invert the beaker over it and allow the flame to go out.
- Wipe the mist that forms in the beaker with a strip of cobalt(II) chloride test paper.
- Repeat the lighting and extinguishing procedure using a 500 mL Erlenmeyer flask.

Take precautions when working near a flame: tie back long hair and keep clothing away from the flame.

- When the flame goes out, lift the flask and turn it upright. Add 10 to 20 mL of limewater to the flask, stopper the flask, and swirl and shake the solution.
- Relight the candle. Holding the beaker upright, lower it until the bottom is about halfway down the flame height. Move the beaker back and forth in the flame for a few seconds, and then remove it.

(a) What product can you detect by condensing it to visible liquid droplets?
(b) What is observed in the cobalt(II) chloride test, and how does this evidence verify your answer to (a)?
(c) What predicted product is not detectable by condensation?
(d) What is observed in the limewater test, and how does this evidence verify the answer to (c)?
(e) What invisible reactant is assumed to be involved in this reaction? What evidence supports this assumption?
(f) What is the black substance on the bottom of the beaker? Why isn't this substance predicted by the reaction equation?
(g) Is the black product a major or minor product? What would you need to know to answer this question?

7.1 Interpreting Chemical Reaction Equations

Science and technology are different activities, but they are mutually interdependent. Either may lead the other in a cycle of expanded knowledge and new abilities to do things. Problems solved by new technology naturally make scientists curious to know and explain why the technology works. The scientific activity that results in new knowledge often results in that knowledge being turned to practical uses by engineers.

Often in our society, and especially in the media, there is a tendency to confuse technology with gadgetry and to assume that the term "technology" just refers to the manufactured devices we use. But there is much more to it than that. Technology also includes the organized processes and skills we develop for doing things. For example, using alphabetical order in a filing system is a technology we use to make it easy to locate the files. In fact, using any alphabet is an incredibly important technology process, just as communicating by human speech is a critical technological skill.

The goal of science is to understand and explain the natural world. In addition, science

- is an *international* discipline
- is involved with *natural* products and processes
- is more *theoretical* in its approach (often based on pure imagination)
- emphasizes *ideas* and *concepts* over practical applications

Scientific concepts and theories are evaluated by how well they *describe, explain,* and *predict* natural phenomena. They do this on the basis of whether the concepts and theories are logical, supported by evidence, consistent with other theories, simple, and testable. The last point is critical, for a concept that cannot be tested, by definition, lies outside the realm of science.

The goal of technology is to provide solutions to practical problems. In addition, technology

- is often more *localized* in use
- is involved with *humanly developed* processes and products
- is more *empirical* in its approach (often based on pure trial-and-error experience)
- emphasizes *methods* and *materials* over understanding

Technological skills, products, and processes are evaluated by how well they *work* to solve practical problems. They do this on the basis of whether the skills, products, and processes are simple, reliable, efficient, sustainable, and economical. A trial-and-error approach can be an effective way to get an unknown system working. It is basically how we learn to walk. The trial-and-error approach is also the system we use to become proficient at a new video game. Anything people use or do to try to make their lives proceed the way they want may properly be termed a technology.

Technologies can be conveniently organized, by scale, into three approximate classes:

- *Industrial technologies* usually involve the very large-scale production of substances from natural raw materials. Examples include mining, oil refining, and the large-scale production of chemicals, such as ammonia and sulfuric acid.
- *Commercial technologies* are medium-scale processes involved in the production of goods at the level of individual business. Examples include the factory production of computers, home appliances, cleaning compounds, and radiator antifreeze.
- *Consumer technologies* involve the use, by individuals on a personal level, of products and processes such as mobile phones, shampoo, shrink-wrap packaging, remote car starters, online banking, and debit cards (**Figure 1**).

DID YOU KNOW?

Technology—Human Nature?

Technological products and processes seem to have always been a part of all human societies. Anthropologists credit the mastery of stone chipping and fire making with starting humans on the road to better controlling and then better understanding their environment. All Aboriginal cultures have developed specialized technology based on solving specific problems. The processes use available materials, and the product addresses a particular need. One example is the birchbark canoe. The need was a light, sturdy craft for an area of many waterways, where land travel is often circuitous and difficult. The birchbark canoe was (and sometimes still is) an elegant solution.

A second universal human factor is imagination—the ability to wonder about things that do not already exist. Wondering "why" and "how" is the key beginning to the knowledge-gathering system we call science.

Figure 1
Debit cards use consumer-scale technology.

Technologies are created and developed to solve social problems. In the process, they often help bring about significant social change. For example, buildings could be no higher than about six stories until the elevator came into use. Shopping malls would be impossible without common access to motor vehicles.

A serious practical question about all new technology is whether the technology will be *sustainable*. There are several reasons why a new technology may not last.

- A technology may become obsolete as a *better technology* replaces it. Computer technology is the classic example: most discarded computers still work perfectly well; they have just been replaced by faster, more powerful models.
- A technology may be based on *a nonrenewable resource*. Over some time period it will either become unavailable or prohibitively expensive. The petroleum industry (the main energy supply for our society) is based on technology that will become useless when we run out of accessible oil (**Figure 2**).
- A technology may produce by-products that *damage the environment*. Unintended harmful effects may well outweigh the usefulness of the technology. Even though the technology still works well for solving the original problem, it will be discontinued. For example, the fluids first used in refrigeration technology were discontinued because of damage to the atmosphere's ozone layer.

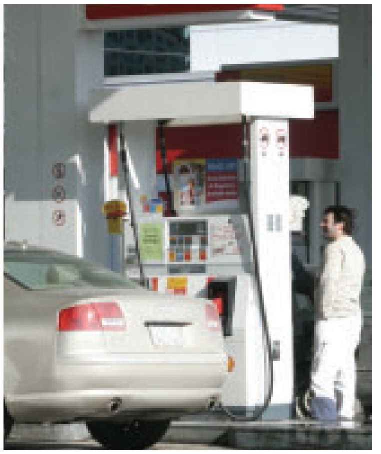

Figure 2
Gasoline as an auto fuel must inevitably run out someday when crude oil reserves are eventually used up.

Of course, nothing lasts indefinitely. If science has taught us anything, it has taught us that everything changes on Earth—it always has, including the shape and position of the continents. The only variation is the time scale involved. We strongly relate time to personal perception. Anything that changes very little over a century, for example, is often regarded as "unchanging," while changes (upgrades) to computer software may seem to happen every week. The point is that we cannot discuss the sustainability of any technology without specifying what we mean and how long a time is involved; otherwise, the discussion necessarily becomes meaningless.

Implementing new technology nearly always has unforeseen effects. Such unintended consequences may be unimportant, or even useful. However, sometimes they create a huge new problem. Use of the pesticide DDT in the mid-1900s decimated many raptor populations, including the Peregrine falcon (**Figure 3**). The subsequent banning of this pesticide has led to a recovery in the numbers of these birds. When society introduces a new technology, we must take extra care and make careful observations.

Figure 3
If a female Peregrine falcon ingests DDT pesticide residue, her eggs would form with shells that are far too thin. The population of these beautiful birds is on the rise today, but they were an endangered species in Canada just a few years ago due to this effect. An international ban on DDT has dramatically increased the survival rate of Peregrine falcon chicks.

Practice

1. List four significant differences between science and technology.
2. Give a clear example of a technology at each of the three levels of scale. Your example should not be one that has already been used in the above text.
3. Classify each of the following questions as to whether it would more likely require a scientific or a technological activity to find an answer. Do not actually answer the questions.
   (a) What coating on a nail will reduce corrosion?
   (b) Which chemical reactions are involved in the corrosion of iron?
   (c) What is the accepted explanation for the chemical formula of water?
   (d) What process produces a continuous thread of nylon?
   (e) Why is a copper(II) sulfate solution blue?
   (f) How can automobiles be designed to make them safer to operate?
4. What does it mean to say a technology must be sustainable? Give an example of any current technology that is likely to be sustainable and another that is certain not to be sustainable. State your reasoning.
5. We often call our society "technological" because our lifestyle depends so heavily on manmade things. Can the lifestyle of Canada's original Aboriginal peoples before 1500 AD be considered nontechnological? Explain your answer, giving examples.

Figure 4
Roberta Bondar (1945–)

WEB Activity

Canadian Achievers—Roberta Bondar

In 1992, Roberta Lynn Bondar (**Figure 4**) became the first Canadian woman in space when she flew on the space shuttle *Discovery* as a payload specialist. Her mission was the first international microgravity laboratory mission.

1. What academic studies did Bondar complete before becoming an astronaut?
2. List three research projects conducted by Bondar.

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CAREER CONNECTION

Aerospace Engineer

Aerospace engineers specialize in all aspects of developing structures and materials related to flight and space travel, including designing aircraft, surveillance systems, satellites, and rockets. Some aerospace engineers test aircraft and spacecraft prototypes experimentally by constructing models to test their performance.

Find out more about aerospace engineering in Canada and the educational requirements for Canada's space program.

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Chemical Reaction Equations

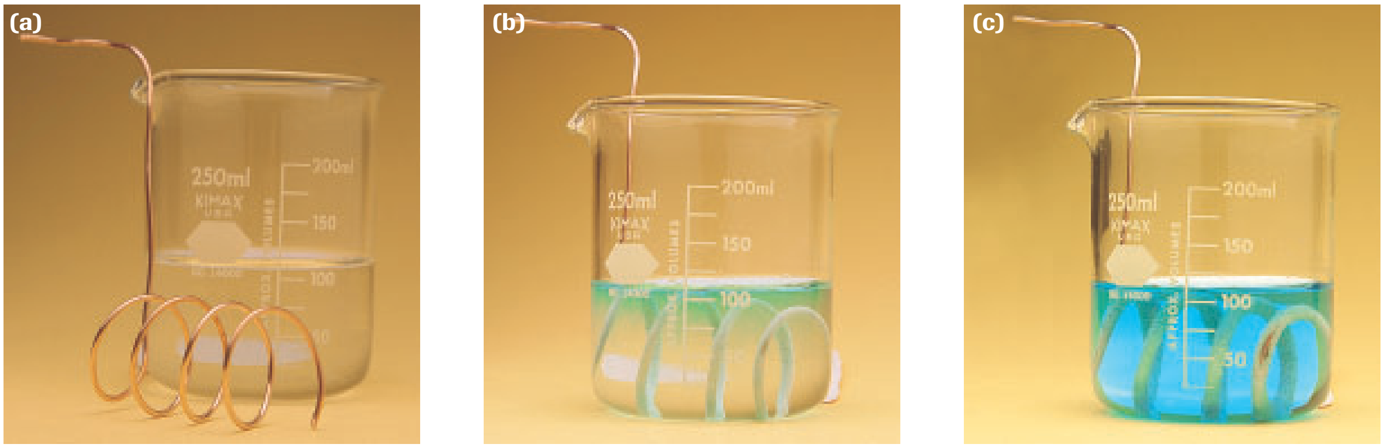
Industrial chemists and chemical engineers must always be concerned about the conditions within and surrounding chemical reactions. After all, a main goal of technology is to develop products and processes that solve practical problems based on criteria such as efficiency, reliability, and cost. Students of chemistry must learn to think of reaction conditions. Chemistry students must also develop the ability to describe, explain, and predict practical outcomes from the scientific knowledge they have acquired. The first question to be asked is, Precisely what do reaction equations tell us? And the next and perhaps even more important question is, What do they not tell us?

Consider the single replacement reaction equation

$$Cu(s) + 2\,AgNO_3(aq) \rightarrow 2\,Ag(s) + Cu(NO_3)_2(aq)$$

How well does this balanced equation describe and explain the reaction? If you were asked to do the reaction, given that you now "know" the balanced chemical equation, you would have to consider several practical questions: What does pure copper look like? What does an aqueous solution of silver nitrate look like? What kind of apparatus should be used to contain this reaction? Does it matter if the silver nitrate is dissolved in water? How much copper would be reasonable to use? How much silver nitrate would be reasonable to use? How much water would be reasonable to use? Some of these questions can be answered by looking at **Figure 5**, which shows photographs taken before, during, and after this particular reaction. However, if you had never seen the reaction or pictures of it, consider how difficult finding the answers might be. For some of the questions, your memory could help; for others, the chemistry concepts you have learned to this point would indicate an answer, or at least a partial answer.

Figure 5
(a) Copper wire and a beaker with aqueous silver nitrate solution
(b) A few moments after the wire is immersed
(c) The beaker contents after 24 h



Limitations of Reaction Equations

If you performed the Exploration activity at the beginning of this chapter, then you detected evidence of an invisible substance. Consider the following similar reaction:

$$C_{25}H_{52}(s) + 38\ O_2(g) \rightarrow 25\ CO_2(g) + 26\ H_2O(g) + \text{heat energy}$$

This reaction equation can represent a typical substance found in candle wax reacting in a burning candle. In this reaction, the only thing you normally see is some light in a special region, the flame. The chemical products are invisible gases, and they are free to mix with other invisible gases (the air). Evidence suggests that the flame is visible for this reason: at one point in the reaction, carbon atoms have broken free of the wax molecule, but they have not yet combined with oxygen. These atoms are very hot, and they glow for an instant, emitting the light you see (**Figure 6**).

Figure 6
A candle flame is easily visible because an intermediate stage of the combustion reaction involves the emission of light.

During winter months in Canada's far north, periods of darkness can last for a very long time. Aboriginal peoples learned long ago how to apply a very simple and effective technology to provide long-burning dependable light from readily available materials. Fat from animals (such as seals, bears, and whales) was rendered (clarified by cooking) to produce a clear oil. A bowl made of some noncombustible material (stone or shell) held the liquid (**Figure 7**), while a wick made of moss was placed at one end of the bowl. With the lower end of the wick submerged in the liquid, and the upper end above the liquid, the moss wick acted to draw the oil up to the flame zone by capillary action.

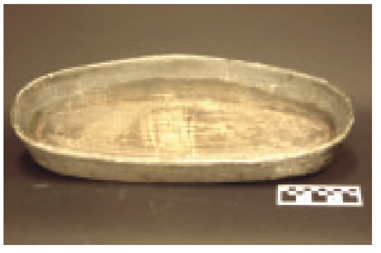


Figure 7
This stone bowl was formed by patiently rubbing with a harder type of stone. When filled with rendered seal oil and with a moss wick added at one end, it makes a good smoke-free light source.

Note that the candle equation above gives "solid" as the state of the wax before it reacts. But wax must be in a gaseous state to burn—just try lighting a candle at the bottom, where there is no wick. The function of a wick is to draw melted wax up by capillary action, moving it into a very hot zone where the wax will vaporize. Once vaporized and mixed with air, it will react (burn) fast enough to set up a continuous reaction.

Chemical reaction equations also do not describe or explain the following:

- A reaction equation usually communicates little about the *pressure and temperature conditions* under which a reaction might occur or might actually be done. This necessary information is sometimes written above the arrow in an equation. To this point, you assume SATP conditions unless told otherwise, but many reactions occur at conditions that are not SATP. For instance, most of the complex reactions that happen inside your body occur at about 37 °C.
- A reaction equation communicates nothing about the *progress and process* of a reaction. It describes what is present before anything happens and what is present after any changes have stopped. It does not describe what actually happens during the reaction or anything about how long the reaction process might take.
- Most importantly, a reaction equation communicates nothing about measurable *quantities* of reactants in any form that you can use directly. An equation is "equal" in chemical symbols because a reaction is "equal" in chemical entities; that is, we believe that entities rearrange and become bonded differently in reactions, but the numbers and kinds of entities do not change. If the numbers of entities do not change, their total mass cannot change. This is the explanation for the law of conservation of mass, which describes a property that can be used directly. However, for most of the questions that arise about quantities of chemicals in reactions, this is not enough.

Equations may be read in terms of single entities (atoms/molecules/ions) or chemical amounts (in moles) of these entities, neither of which describes any property you can see, count, or measure directly in a laboratory or classroom. For instance, if you wanted to make 10.0 g of silver with the copper–silver nitrate reaction, a new set of questions would arise. This is the primary subject of this chapter.

Reaction Assumptions

Making assumptions is common, and not just in chemistry. We do it all the time; it's the way that the human mind works. The danger is that you might be unaware that an assumption is being made. This is like crossing the street without looking, assuming subconsciously that all the drivers on the street will obey crosswalk rules. It's a good idea to look anyway! In chemistry, you must be aware of the major assumptions normally made for reactions, so you will notice when an exception occurs. A reaction equation always carries with it many assumptions, untested statements considered to be correct without proof or demonstration. These must be known and understood to make practical predictions in many cases, especially in this unit. To this point in your study of chemistry, you assume the following:

Figure 8
A spectacular example of combustion of a metal is the burning of steel wool (iron) in oxygen.

- Reactions are *spontaneous*. Reactions will occur when the reactants are mixed for all examples you are given. However, you have not yet studied any generalization, law, or theory that will let you predict whether given substances will react.
- Reactions are *fast*. For a reaction to be useful, either in a laboratory or in industry, the reaction must occur within a reasonable time. Consider the following simple balanced reaction equation, which gives no indication about the rate of the reaction:

$$4\ Fe(s)\ +\ 3\ O_2(g)\ \rightarrow\ 2\ Fe_2O_3(s)\ +\ \text{heat energy}$$

If this reaction represents fine strands of steel wool ignited in pure oxygen, it is over in a few seconds (**Figure 8**). If it represents the fine iron powder in a hand warmer pouch reacting in air, it will take several hours (**Figure 9**). If the reaction represents the eventual corrosion of an automobile, it could take many decades

Figure 9
A hand warmer pouch is sold sealed in an airtight plastic wrap. When the plastic outer wrap is removed, air filters in slowly through the porous pouch, and the reaction of the powdered iron (and other solids) inside the pouch can occur. The pouch is a carefully designed technological device, made to get the reaction to release heat at a rate that will warm your hand but not burn a hole in your coat pocket. One pouch has been cut open to show the powdered solid reagents inside.

(**Figure 10**). Controlling the rate of chemical reactions is extremely important; it concerns every aspect of society, from the speed at which fuel burns in a car, to the time dental epoxy takes to set and bond a tooth, to the speed of all the reactions that make up human metabolism.

- Reactions are *quantitative*. A **quantitative reaction** is one that is more than 99% complete; in other words, at least one reactant is essentially completely used up. Another way of saying this is that the reaction goes to completion.
- Reactions are *stoichiometric*. A **stoichiometric reaction** means that there is a simple, whole-number ratio of chemical amounts of reactants and products. In other words, the coefficients that you predict for a balanced chemical equation do not change when the reaction is repeated several times, even under different conditions.

These four key assumptions—that chemical reactions are spontaneous, fast, quantitative, and stoichiometric—will be tested and evaluated later in your chemistry education. These assumptions are particularly important when you do quantitative studies of chemical reactions in the rest of this unit.

Figure 10
When solid sheet metal is protected by paint, it can take many years to reach this level of oxidation (rusting). Spread over this time span, the heat energy released is completely undetectable.

Practice

6. What does a balanced reaction equation directly communicate?
7. State three important aspects of a chemical reaction that are not communicated by the balanced chemical equation.
8. List the four major assumptions usually made about chemical reactions.
9. List three criteria that are often used to evaluate a technology.

Net Ionic Equations

Now we return to the original reaction equation example:

$$Cu(s) + 2\,AgNO_3(aq) \rightarrow 2\,Ag(s) + Cu(NO_3)_2(aq)$$

Collision–reaction theory is useful to consider when a solution is involved in a reaction. Many reactions will *only* occur in solution. Dissolving a reactant is often the only easy way to get its entities separated from each other so they can collide with entities of another reactant. To do this reaction, we could consider vaporizing the copper to separate its atoms. However, the temperature required is over 2500 °C, and the vapour is highly toxic and highly reactive. On the other hand, all the silver and nitrate ions in solid silver nitrate can be separated just by placing a sample in water and stirring it; an ordinary open glass beaker will work perfectly well as the reaction container. As a side benefit, since higher concentrations of substances react faster (more collisions), we can often easily control the rate of a reaction in solution.

One drawback, of course, is that often there is no direct visible way to tell a solution apart from the pure solvent. Silver nitrate solution looks exactly like pure water. In **Figure 5(a)** (page 278), the solid silver nitrate was already dissolved before the picture was taken, so you could not tell what the pure compound looked like. A few solutes produce visibly coloured solutions (like copper(II) compounds), and these sometimes allow visual identification of substances (**Figure 11**). However, usually chemical analysis techniques are needed to find out what, and how much, solute is in an aqueous solution. It is tempting to say that because the solution turns blue and copper(II) nitrate is produced in solution, that obviously the copper(II) nitrate is blue, but there is more to it than that. To really understand how dissolved ionic compounds react, it is necessary to write the equations in a form that more correctly represents the actual state of the entities present.

Figure 11
The yellow and orange colours seen here are characteristic of aqueous chromate and dichromate ions, respectively.

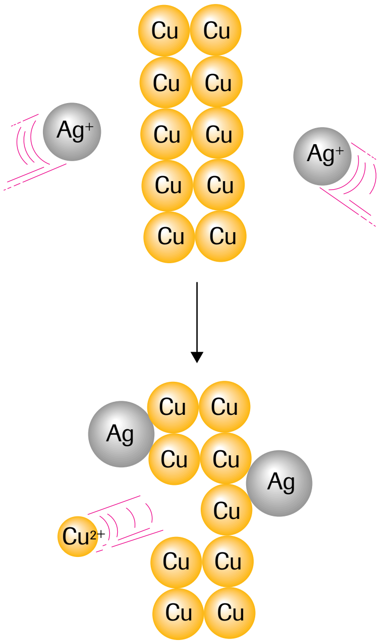


Figure 12
A model of the reaction of copper metal and silver nitrate solution illustrates aqueous silver ions reacting at the surface of a solid copper strip.

For writing net ionic equations, it is useful to refer to Table 3 in Section 5.2. The table indicates that a highly soluble ionic compound dissociates into individual ions as it dissolves in solution. If we rewrite the chemical equation to show that entities of the dissolved ionic compounds are actually present as separate (dissociated) aqueous ions, it then looks like the following:

$$Cu(s) + 2\,Ag^+(aq) + 2\,NO_3^-(aq) \rightarrow 2\,Ag(s) + Cu^{2+}(aq) + 2\,NO_3^-(aq)$$

Notice that when written this way, the equation makes it obvious that the nitrate ions do not change at all in the course of the reaction. They are like the beaker and the water in that they are not part of the reaction itself; they just help create an environment where the reaction can occur. If we write the equation again, leaving out any entities that do not change in the reaction, the result is the **net ionic equation**:

$$Cu(s) + 2\,Ag^+(aq) \rightarrow 2\,Ag(s) + Cu^{2+}(aq)$$

The equation can now be interpreted as follows: "If solid copper is placed in an aqueous solution of silver ions, solid silver will form, and copper(II) ions will form in solution" (**Figure 12**). This reaction statement looks different because it does not specify *what* silver compound was used to make the silver ion solution. How important is this point?

SAMPLE problem 7.1

A student mixed solutions of lead(II) nitrate and sodium iodide and observed the formation of a bright yellow precipitate. Another student recorded the same observation after mixing solutions of lead(II) acetate and magnesium iodide. Are these different reactions?

The balanced chemical equations for these two double replacement reactions show some similarities and some differences.

$$Pb(NO_3)_2(aq) + 2\,NaI(aq) \rightarrow PbI_2(s) + 2\,NaNO_3(aq) \quad (1)$$

$$Pb(CH_3COO)_2(aq) + MgI_2(aq) \rightarrow PbI_2(s) + Mg(CH_3COO)_2(aq) \quad (2)$$

Using Arrhenius' theory of dissociation, these reactions can be described more precisely. Each of the highly soluble ionic compounds is believed to exist in aqueous solution as separate ions. For reaction (1),

$$Pb^{2+}(aq) + 2\,NO_3^-(aq) + 2\,Na^+(aq) + 2\,I^-(aq) \rightarrow PbI_2(s) + 2\,Na^+(aq) + 2\,NO_3^-(aq)$$

It is apparent that some reactant ions—sodium and nitrate ions—are unchanged in this reaction. Ignoring these ions, you can write a net ionic equation, which shows only the entities that change in a chemical reaction:

$$Pb^{2+}(aq) + \cancel{2\,NO_3^-(aq)} + \cancel{2\,Na^+(aq)} + 2\,I^-(aq) \rightarrow PbI_2(s) + \cancel{2\,Na^+(aq)} + \cancel{2\,NO_3^-(aq)}$$

$$Pb^{2+}(aq) + 2\,I^-(aq) \rightarrow PbI_2(s) \text{ (net ionic equation)}$$

Applying the same procedure to reaction (2):

$$Pb^{2+}(aq) + \cancel{2\,CH_3COO(aq)} + \cancel{Mg^{2+}(aq)} + 2\,I^-(aq) \rightarrow$$

$$PbI_2(s) + \cancel{Mg^{2+}(aq)} + \cancel{2\,CH_3COO(aq)}$$

$$Pb^{2+}(aq) + 2\,I^-(aq) \rightarrow PbI_2(s) \text{ (net ionic equation)}$$

The net ionic equations are identical for reactions (1) and (2), as are the observations (**Figure 13**). We can therefore say that the reactions are the same.

Figure 13
Solid, yellow lead(II) iodide precipitates when any lead(II) ion solution is mixed with any iodide ion solution. Reaction mixtures (1) and (2) from Sample Problem 7.1 produce exactly the same visible evidence.

Ions that are present but do not take part in (change during) a reaction are called **spectator ions**. These ions can be likened to spectators at a sports event; the spectators are present but are not part of the game. The conclusion to be drawn from the net ionic equations in Sample Problem 7.1 is that there were not two different reactions, just two different sets of substances used to make the same reaction occur.

COMMUNICATION *example* 1

Write the net ionic equation for the reaction of aqueous barium chloride and aqueous sodium sulfate. Refer to Section 5.5 and to the solubility table on the inside back cover of this textbook.

Solution

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2\,NaCl(aq)$$

$$Ba^{2+}(aq) + \cancel{2\,Cl^-(aq)} + \cancel{2\,Na^+(aq)} + SO_4^{\;2-}(aq) \rightarrow BaSO_4(s) + \cancel{2\,Na^+(aq)} + \cancel{2\,Cl^-(aq)}$$

$$Ba^{2+}(aq) + SO_4^{\;2-}(aq) \rightarrow BaSO_4(s)$$

Learning *Tip*

When eliminating or cancelling spectator ions, they must be identical in every way: chemical amount, form (atom/ion/molecule), and state of matter. Occasionally, the amount may be different while the form and state are identical. In this case, you may only cancel equal amounts.

Net ionic equations are useful in communicating reactions other than double replacement reactions. Communication Example 2 is a good illustration.

COMMUNICATION *example* 2

Write the net ionic equation for the reaction of zinc metal and aqueous copper(II) sulfate.

Solution

$$Zn(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4(aq)$$

$$Zn(s) + Cu^{2+}(aq) + \cancel{SO_4^{\;2-}(aq)} \rightarrow Cu(s) + Zn^{2+}(aq) + \cancel{SO_4^{\;2-}(aq)}$$

$$Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$$

COMMUNICATION *example* 3

Write the net ionic equation for the reaction of hydrochloric acid and barium hydroxide solution.

Solution

$$2\,HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2\,H_2O(l)$$

$$2\,H^+(aq) + \cancel{2\,Cl^-(aq)} + \cancel{Ba^{2+}(aq)} + 2\,OH^-(aq) \rightarrow \cancel{Ba^{2+}(aq)} + \cancel{2\,Cl^-(aq)} + 2\,H_2O(l)$$

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \text{ (coefficients reduced to 1)}$$

Learning *Tip*

Evidence indicates that protons in solution really exist attached to one or more water molecules. However, for writing ordinary net ionic equations for reactions involving aqueous strong acids, the entity symbol used is usually $H^+(aq)$, and not $H_3O^+(aq)$. This is just a matter of convenience; the H^+ symbol is quicker and easier to write. (Note this usage in Communication Example 3.)

In Unit 8 you will study situations where use of the hydronium ion symbolism is necessary and more useful.

SUMMARY *Writing Net Ionic Equations*

Step 1: Write a complete balanced chemical equation.
Step 2: Dissociate all high-solubility ionic compounds and ionize all strong acids to show the complete ionic equation.
Step 3: Cancel identical entities that appear on both reactant and product sides.
Step 4: Write the net ionic equation, reducing coefficients if necessary.

Figure 14
A water treatment facility

Practice

10. In a laboratory test of the metal activity series, a student places a strip of lead metal into aqueous silver nitrate. Write the net ionic equation for the reaction that occurs.

11. (a) In a water treatment facility (**Figure 14**), sodium phosphate is added to remove calcium ions from the water. Write the net ionic equation for the reaction of aqueous calcium chloride and aqueous sodium phosphate.
(b) Identify the spectator ions in this reaction.

12. Some natural waters contain iron ions that affect the taste of the water and cause rust stains. Aeration converts any iron(II) ions into iron(III) ions. A basic solution (containing hydroxide ions) is added to produce a precipitate. Write the net ionic equation for the reaction of aqueous iron(III) ions and aqueous hydroxide ions.

13. A nitric acid spill is quickly neutralized by pouring a sodium hydrogen carbonate (baking soda) solution on it. Write the chemical equation and the net ionic equation for this neutralization reaction. Identify the spectator ions by name.

14. When you open a can of pop, the pressure inside the can is released. This allows the aqueous carbonic acid to decompose, forming carbon dioxide gas and water.
(a) Write the net ionic equation for this reaction.
(b) Write a statement about the dual role of water molecules in this particular reaction.

Limiting and Excess Reagents

Revisiting the questions at the beginning of this section raises more points of interest. Consider the reaction demonstration in **Figure 5** (page 278) again. What is in the container after the reaction is finished? You can directly observe evidence for the silver, you know that the blue colour is likely from the copper(II) nitrate, and you can see that there is still a lot of solid copper left over. Of course, there is also a lot of water present. When no further changes appear to be occurring, we assume that all the silver nitrate that was initially present has now been completely reacted.

Is the silver nitrate really all gone? It is invisible in this system, so how can you tell? For reactions in which we care about quantities of substances involved, making sure that a measured reagent reacts completely becomes critically important. The standard method for this is to ensure that the measured reactant is a limiting reagent. A **limiting reagent** is the reactant whose entities are completely consumed in a reaction, meaning the reaction stops when—and because—all of this reactant is used up and none remains. To make sure that this happens, more of the other reactant must be present than is required for the reaction; otherwise, you would run out of it first. A greater quantity of this reactant than is necessary is deliberately added to the reaction system, and it is described as an excess reagent. An **excess reagent** is the reactant whose entities are present in surplus amounts, so that some remain after the reaction ends. In our reaction example, much more copper was used than was needed, as evidenced by the unreacted copper, so copper is the excess reagent. We assume that the reaction ended when there were no more silver ions left to react, so silver nitrate was the limiting reagent.

Most of our unanswered original questions are about "how much." After all, people do chemical reactions for specific reasons, and the activity nearly always involves knowing, measuring, or predicting quantities of something. The rest of this chapter is about combining many of the concepts you have learned so far to identify, calculate, and predict quantities of chemicals involved in reactions.

DID YOU KNOW ?

Diagnostic Tests for Ions
You can easily create diagnostic tests for many specific ions using the solubility chart to find an oppositely charged ion that would produce a low solubility product. For example, here is a method to test for silver ions in a solution: Add a few drops of aqueous sodium chloride to the solution. If a precipitate forms, then silver ions are still present in the solution.

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CAREER CONNECTION

Chemical Engineer
Chemical engineers are employed in a wide range of manufacturing and processing industries, consulting firms, government, research, and educational institutions. Among other duties, chemical engineers develop chemical processes in which reactions must be known.

What education is required to become a chemical engineer?

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Section 7.1 Questions

1. Most Albertans use natural gas to heat their homes and to produce hot water. After refining, natural gas is composed almost entirely of methane.
(a) Write the balanced chemical equation for the complete combustion of methane.
(b) What specific information is given directly by this chemical equation?
(c) What are some things that you do not know about this reaction?

2. At this stage in your chemistry education, you need to assume that chemical equations represent reactions that are spontaneous, fast, quantitative, and stoichiometric.
(a) In your own words, explain what each of these assumptions means.
(b) In general terms, explain how you would check these assumptions for a particular reaction.

3. An acceptable method for the treatment of soluble lead waste is to precipitate the lead as a low solubility lead(II) silicate.
(a) Write the net ionic equation for the reaction of aqueous lead(II) nitrate and aqueous sodium silicate.
(b) What can we assume about the ambient conditions and the container that likely could be used?
(c) Identify the spectator ions in this reaction.

4. Bromine is a disinfectant commonly used in swimming pools. One industrial method of producing bromine is to react sea water, containing sodium bromide, with chlorine gas. Write the net ionic equation for this reaction.

5. (a) Strontium compounds are often used in flares because their flame colour is bright red (**Figure 15**). One commercial example of the production of strontium compounds is the reaction of aqueous solutions of strontium nitrate and sodium carbonate. Write the net ionic equation for this reaction.
(b) Suggest another compound in solution (other than sodium carbonate) that would react with strontium nitrate solution to produce a reaction with the same net ionic equation as the reaction in (a).

Figure 15
The bright red of the flare is easily visible to passing motorists.

6. In a hard water analysis, sodium oxalate solution reacts with calcium hydrogen carbonate present in the hard water to precipitate a calcium compound. Write the net ionic equation for this reaction.

7. Write a net ionic equation for the reaction of vinegar (acetic acid solution) with a scale deposit in a kettle (assume solid calcium hydroxide).

8. State why it is desirable, in a quantitative chemical analysis of a substance, to use an excess of one reactant.

9. For a particular reaction, how are the interests of a research chemist different from the interests of an industrial chemist or engineer? How are they similar?

10. Introduction of new technology often has unintended consequences. It is not unusual for these consequences to be beneficial. One example is the discovery of using lasers to transfer information to and from DVDs. Unintended consequences can be very undesirable, however. Write statements including the terms "obsolete" and "recycle" to describe the current social problem that is an undesirable consequence of the rapid development of personal computer technology. Use an ecological point of view.

11. Mobile (cell) phone cameras (**Figure 16**) are a technology that is rapidly expanding in use worldwide. Write one statement about some aspect of this technology from each of the following perspectives:
(a) economic
(b) societal
(c) ethical
(d) environmental

Figure 16
Many cellphones now come with a camera, which is raising new issues regarding privacy.

12. State what measurements would normally have to be taken in a lab to allow you to calculate the chemical amount in a sample of each of the following substances:
(a) $CH_4(g)$
(b) $NaCl(s)$
(c) $C_6H_6(l)$
(d) 6.0 mol/L $HCl(aq)$

7.2 Gravimetric Stoichiometry

CAREER CONNECTION

Chemical Technologist
Chemical technologists and technicians find employment in the private sector as well as in government. They perform chemical tests and help create procedures in the laboratory, from routine processes to more intricate procedures needed in complex research projects.

What level of education is required to become a chemical technologist?

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Chemical engineers and technologists design and control chemical technology in a processing plant. Like all technology, the goal is to solve practical problems, such as producing fertilizers for the agricultural industry, improving the combustion qualities and environmental impacts of fuels, and creating better and safer water treatment processes. Typically, society provides the practical problem to be solved, science provides some or all of the understanding, and technology is developed to come up with the solution. The water purification unit used by Canada's Disaster Assistance Relief Team (DART) is a good example of this process (**Figure 1**(**a**)). The problem is the need for pure water in remote locations where only contaminated water is available; various sciences such as chemistry, biology, and physics provide the empirical and theoretical knowledge; and engineers and technologists create the technology to provide the solution to the original problem.

Typically, any such situation actually involves a whole series of problems, each requiring a specific technology to achieve the desired results. For example, the DART system produces a huge volume of water, which is temporarily stored in containers that must be very big, yet easily portable. As a result of polymer science, we now have high-strength synthetic plastics. Technologists used this knowledge to create collapsible water storage bladders for the DART team (**Figure 1**(**b**)). Then, of course, the water has to be pumped out of the storage bladders somehow, and that means energy must be provided to run the pumps, and on and on the process goes.

Most chemical technologies require quantitative predictions of raw materials and products. Quantitative predictions made to ensure that a commercial or industrial process works well are based largely on an understanding of the relative quantities of

(a)

(b)

Figure 1
(a) This rugged water treatment unit was used by the Canadian Armed Forces DART unit in the Asian tsunami relief effort in Sri Lanka in December 2004. The unit can produce 120 kL (120 m^3) of water per day from any water source, including chemically contaminated water. A large diesel engine runs powerful pumps that push water through large bundles of tubular membranes. Each individual tiny tube has many submicroscopic holes; the holes are so small that only water molecules can get through, so pure water emerges.
(b) Sergeant Shane Stachnick, from Rosetta, Alberta, distributes water from the collapsible storage bladder, which can hold 10 000 L (10 t), to the local people.

reactants consumed and products produced in a chemical system. This understanding can be entirely empirical, determined by trial and error, but more often, it is related to a knowledge of the balanced chemical reaction equation. For all chemical reactions where quantities are important—whether in industry, commerce, research, or analysis—a balanced equation is necessary because it describes the reaction stoichiometry. The **stoichiometry** of a reaction is the description of the relative quantities of the reactants and products by chemical amount, that is, in moles. Any prediction or calculation from any measured quantity of any substance in a reaction must necessarily be based on the stoichiometry of the reaction. Therefore, as you recall from the previous section, reactions must be stoichiometric, but we also assume that the reactions will be spontaneous, fast, and quantitative.

DID YOU KNOW ?

Meaning of Stoichiometry

Stoichiometry (stoy-kee-**ah**-meh-tree) is derived from the Greek words *stoicheion* (element) and *metron* (measure).

INVESTIGATION 7.1 *Introduction*

Decomposing Malachite

Copper(II) hydroxide carbonate, commonly called basic copper carbonate and also known as malachite, is a double salt with the chemical formula $Cu(OH)_2 \cdot CuCO_3(s)$. This double salt decomposes completely when heated to 200 °C, forming copper(II) oxide, carbon dioxide, and water vapour. Complete the Prediction using the balanced chemical equation. Include safety and disposal steps in your Procedure. Organize the data and create the graph using suitable software.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
○ Hypothesis	● Procedure	
● Prediction	● Evidence	

Purpose

The purpose of this investigation is to test the assumption that a chemical reaction is stoichiometric.

Problem

How is the chemical amount of copper(II) oxide produced related to the chemical amount of malachite reacted in the decomposition of malachite?

Design

A known mass of malachite (manipulated variable) is heated strongly until the colour changes completely from green to black. The mass of black copper(II) oxide (responding variable) is determined. The results from several laboratory groups are combined in a graph to determine the ratio of chemical amounts.

To perform this investigation, turn to page 304.

Practice

1. What is the main goal of technology? Illustrate this with one example.
2. Most modern automobiles have improved fuel economy and produce less pollution compared to those built a number of years ago.
   (a) What role did science likely play in these technological developments?
   (b) What role did society likely play?
3. Technologies can be classified according to the scale of the technology. Identify the three scales, and write a brief description of each.
4. In your own words, explain what stoichiometry means. How have you been using this concept almost since you started studying chemistry?
5. Which of the four assumptions about chemical reactions is tested in Investigation 7.1? Was this assumption shown to be valid?
6. Using the balanced chemical equation for the decomposition of malachite, how does the chemical amount of copper(II) oxide product compare with the chemical amount of the carbon dioxide product?
7. What evidence do you have that the reaction in Investigation 7.1 was likely quantitative, that is, went to completion?

Calculating Masses Involved in Chemical Reactions

Analysis of the evidence from Investigation 7.1 indicates that when malachite is decomposed, the ratio of the chemical amounts of copper(II) oxide and malachite is a simple mole ratio of 2:1. This is the same ratio given by the coefficients of these substances in the balanced chemical equation. Two moles of copper(II) oxide are produced for each (one) mole of malachite that reacts.

$$Cu(OH)_2{\cdot}CuCO_3(s) \rightarrow 2\,CuO(s) + CO_2(g) + H_2O(g)$$

Unfortunately, there is no instrument that measures amounts in moles directly. A measurable quantity such as mass is required, from which we can predict and analyze the quantities of reactants and products in a chemical reaction. However, the relationship between two substances in a chemical reaction is represented by the mole ratio from the balanced chemical equation.

The procedure for calculating the masses of reactants or products in a chemical reaction is called **gravimetric stoichiometry**. Gravimetric stoichiometry is restricted to chemical amount calculations from mass (gravity) measurement, so the measured substance has to be a pure solid or liquid. However, the calculated mass can be for any other substance in the reaction.

Gas stoichiometry, which you will study in Section 7.3, requires that volume, temperature, and pressure all be considered to calculate the chemical amount. This is because the entities are widely separated from each other and must be held in a sealed container.

Solution stoichiometry, which is covered in Section 7.4, also involves entities that are widely separated, but only the amount concentration and solution volume usually need to be measured to calculate the chemical amount.

SAMPLE problem 7.2

If you decompose 1.00 g of malachite, what mass of copper(II) oxide would be formed?

First, write the balanced chemical equation. Underneath the balanced equation, write the mass that is given (measured) and the symbol *m* for the mass to be calculated, along with the conversion factors. In this example, one mass is given, and the conversion factors (the molar masses) are calculated from the chemical formulas and the information in the periodic table:

$$Cu(OH)_2{\cdot}CuCO_3(s) \rightarrow 2\,CuO(s) + CO_2(g) + H_2O(g)$$

$Cu(OH)_2{\cdot}CuCO_3(s)$	$2\,CuO(s)$
1.00 g	*m*
221.13 g/mol	79.55 g/mol

Second, convert the measured mass of malachite to its chemical amount:

$$n_{Cu(OH)_2{\cdot}CuCO_3} = 1.00\ \cancel{g} \times \frac{1\ mol}{221.13\ \cancel{g}}$$

$$= 0.004\ 52\ mol$$

Third, calculate, using the mole ratio from the balanced equation, the amount of copper(II) oxide that will be produced:

$$\frac{n_{CuO}}{n_{Cu(OH)_2{\cdot}CuCO_3}} = \frac{2}{1}$$

$$\frac{n_{CuO}}{0.004\ 52\ mol} = \frac{2}{1}$$

Learning Tip

In all stoichiometric calculations, the third step is the same. The mole ratio is always used with the coefficient for the unknown (or required) substance as the numerator and the coefficient for the measured (or given) substance as the denominator. Where the single calculation method is shown in this textbook, formulas are written in with the mole ratio and the molar masses and then cancelled to make it clear how the calculation must be correctly set up.

$$n_{CuO} = 0.004\ 52\ \text{mol} \times \frac{2}{1}$$
$$= 0.009\ 04\ \text{mol}$$

Fourth, calculate the mass represented by this amount of CuO:

$$m_{CuO} = 0.009\ 04\ \cancel{\text{mol}} \times \frac{79.55\ \text{g}}{1\ \cancel{\text{mol}}}$$
$$= 0.719\ \text{g}$$

Alternatively, all three steps of such a calculation can be expressed as a single "chained" calculation. When using this method, it is customary to label each quantity and conversion factor and to cancel quantities and labels carefully. The purpose is to keep track of the substances involved.

$$m_{CuO} = 1.00\ \cancel{\text{g Cu(OH)}_2\cdot\text{CuCO}_3} \times \frac{1\ \cancel{\text{mol Cu(OH)}_2\cdot\text{CuCO}_3}}{221.13\ \text{g}\ \cancel{\text{Cu(OH)}_2\cdot\text{CuCO}_3}} \times \frac{2\ \cancel{\text{mol CuO}}}{1\ \cancel{\text{mol Cu(OH)}_2\cdot\text{CuCO}_3}} \times \frac{79.55\ \text{g CuO}}{1\ \cancel{\text{mol CuO}}}$$
$$= 0.791\ \text{g CuO}$$

The certainty of the answer, three significant digits, is determined by the least certain value used in the calculation, 1.00 g. Note that the mass of copper(II) oxide has been obtained by knowing only the balanced chemical equation and the molar masses. No actual experiment was necessary. The following example illustrates how to communicate the stoichiometric method.

COMMUNICATION example

Iron is the most widely used metal in North America (**Figure 2**). It may be produced by the reaction of iron(III) oxide, from iron ore, with carbon monoxide to produce iron metal and carbon dioxide. What mass of iron(III) oxide is required to produce 100.0 g of iron?

Solution

$Fe_2O_3(s)$	+	$3\ CO(g)$	→	$2\ Fe(s)$	+	$3\ CO_2(g)$
m				100.0 g		
159.70 g/mol				55.85 g/mol		

$$n_{Fe} = 100.0\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{55.85\ \cancel{\text{g}}}$$
$$= 1.791\ \text{mol}$$

$$n_{Fe_2O_3} = 1.791\ \text{mol} \times \frac{1}{2}$$
$$= 0.8953\ \text{mol}$$

$$m_{Fe_2O_3} = 0.8953\ \cancel{\text{mol}} \times \frac{159.70\ \text{g}}{1\ \cancel{\text{mol}}}$$
$$= 143.0\ \text{g}$$

or

$$m_{Fe_2O_3} = 100.0\ \cancel{\text{g Fe}} \times \frac{1\ \cancel{\text{mol Fe}}}{55.85\ \cancel{\text{g Fe}}} \times \frac{1\ \cancel{\text{mol Fe}_2\text{O}_3}}{2\ \cancel{\text{mol Fe}}} \times \frac{159.70\ \text{g Fe}_2\text{O}_3}{1\ \cancel{\text{mol Fe}_2\text{O}_3}}$$
$$= 143.0\ \text{g Fe}_2\text{O}_3$$

According to gravimetric stoichiometry, 143.0 g of iron(III) oxide is needed to produce 100.0 g of iron.

Learning Tip

Remember to keep the unrounded values in your calculator for further calculation until the final answer is reported. The values for intermediate calculation are rounded when written down. Follow the calculation process for the Sample Problems on your calculator to review how to do this.

Figure 2
Wrought iron is a very pure form of iron. The ornate gates on Parliament Hill in Ottawa are made of wrought iron. The metal is relatively soft and easily bent into decorative shapes. Wrought iron is also quite corrosion-resistant. When carbon is present in iron in small quantities, the metal becomes much harder and is called steel.

SUMMARY Gravimetric Stoichiometry

Step 1: Write a balanced chemical reaction equation, and list the measured mass, the unknown quantity (mass) symbol *m*, and conversion factors (the molar masses).
Step 2: Convert the mass of measured substance to its chemical amount.
Step 3: Calculate the chemical amount of required substance using the mole ratio from the balanced chemical equation.
Step 4: Convert the chemical amount of required substance to its mass.

Stoichiometry Calculations

measured quantity
solids/liquids $m \rightarrow n$

↓ mole ratio

required quantity
solids/liquids $m \leftarrow n$

Practice

8. Why is a balanced chemical equation necessary when doing a stoichiometry calculation?
9. Powdered zinc metal reacts violently with sulfur (S_8) when heated to produce zinc sulfide (**Figure 3**). Predict the mass of sulfur required to react with 25 g of zinc.
10. Bauxite ore contains aluminium oxide, which is decomposed using electricity to produce aluminium metal (**Figure 4**). What mass of aluminium metal can be produced from 125 g of aluminium oxide?
11. Determine the mass of oxygen required to completely burn 10.0 g of propane.
12. Calculate the mass of lead(II) chloride precipitate produced when 2.57 g of sodium chloride in solution reacts in a double replacement reaction with excess aqueous lead(II) nitrate.
13. Predict the mass of hydrogen gas produced when 2.73 g of aluminium reacts in a single replacement reaction with excess sulfuric acid.
14. What mass of copper(II) hydroxide precipitate is produced by the reaction in solution of 2.67 g of potassium hydroxide with excess aqueous copper(II) nitrate?

Figure 3
The reaction of powdered zinc and sulfur is rapid and highly exothermic. Because of the numerous safety precautions that would be necessary, the reaction is not usually carried out in school laboratories.

DID YOU KNOW?

Refining Aluminium
Aluminium is the most abundant metal in Earth's crust, but it occurs only in chemical compounds such as aluminium oxide, Al_2O_3, the principal constituent of bauxite. Canada has little bauxite but has abundant hydroelectric power. Aluminium oxide imported principally from Jamaica and Australia is refined at Alcan's aluminium refinery in Kitimat, British Columbia, and then exported worldwide. The refinery uses just over 500 kt of aluminium oxide annually to produce about 272 kt of aluminium. To produce this amount of aluminium requires 896 MW of electricity. The B.C. refinery came at a social cost though, with many Aboriginal groups displaced from their homes to allow the flooding necessary for the power dams.

Figure 4
An aluminium refinery

Testing the Stoichiometric Method

The most rigorous test of any scientific concept is whether it can be used to make predictions. If the prediction is shown to be valid, then the concept is judged to be acceptable. The prediction is falsified if the percent difference between the actual and the predicted values is considered to be too great, for example, more than 10%. The concept may then be judged unacceptable. (See "Evaluation" in Appendix B.2.) Percent difference between an experimental value and a predicted value is the primary criterion for the evaluation of an accepted value (such as a constant) or an accepted method (such as stoichiometry). It is assumed that reagents are pure and skills are adequate for the experiment that is conducted.

Filtration (see Appendix C.4) is a common technique used in experimental designs for testing stoichiometric predictions. Stoichiometry is used to predict the mass of precipitate that will be produced, and filtration is used to separate the mass of precipitate actually produced in a reaction (**Figure 5**). Lab Exercise 7.A and Investigation 7.2 test the validity of the stoichiometric method. In all examples, an excess of one reactant is used to ensure complete reaction of the limiting (measured or tested) reagent.

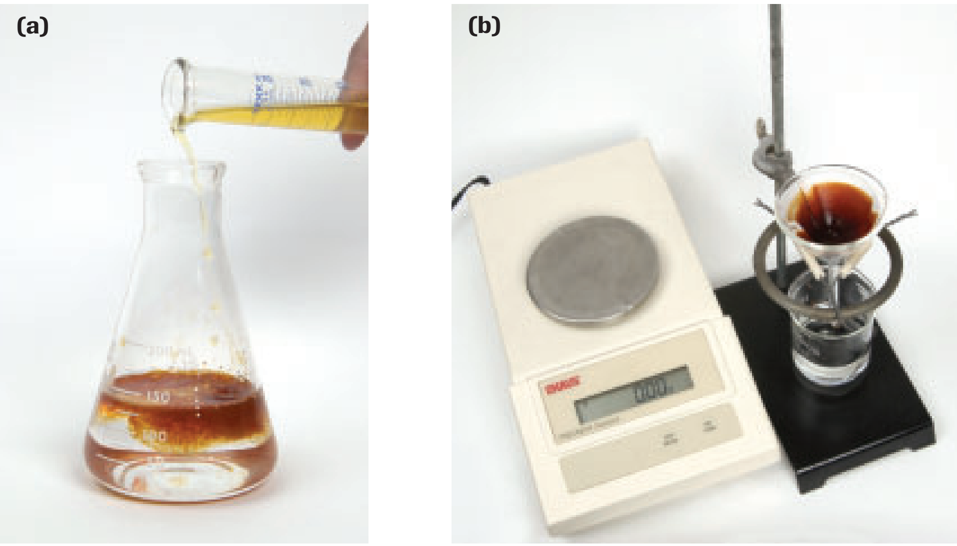

Figure 5
(a) A dissolved substance or ion can often be precipitated out of solution.
(b) A precipitate is filtered and dried, and its mass is measured to determine the amount of substance that was dissolved in the original solution.
The technique of filtration is explained in Appendix C.4, and demonstrated in the video.

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LAB EXERCISE 7.A

Testing the Stoichiometric Method

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ● Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ○ Evidence
- ● Analysis
- ● Evaluation (2, 3)

Purpose

The purpose of this investigation is to test the stoichiometric method. In your evaluation, assume the experiment was valid and that suitable quality evidence was obtained.

Problem

What mass of lead is produced by the reaction of 2.13 g of zinc with an excess of lead(II) nitrate in solution (**Figure 6**)?

Design

A known mass of zinc is placed in a beaker with an excess of lead(II) nitrate solution. The lead produced in the reaction is separated by filtration and dried. The mass of the lead is determined.

Evidence

In the beaker, crystals of a shiny black solid were produced, and all the zinc disappeared.
mass of filter paper = 0.92 g
mass of dried filter paper plus lead = 7.60 g

Figure 6
Zinc reacts with a solution of lead(II) nitrate.

INVESTIGATION 7.2 *Introduction*

Gravimetric Stoichiometry

In this investigation, you will use gravimetric stoichiometry to investigate the reaction of strontium nitrate with excess copper(II) sulfate in an aqueous solution. Use 2.00 g of strontium nitrate and about 3.5 g of copper(II) sulfate–water (1/5), initially dissolving each chemical in about 75 mL of water. Be sure to include safety and disposal instructions in your Procedure.

Purpose

The purpose of this investigation is to test the stoichiometric method.

Report Checklist

○ Purpose	● Design	● Analysis
○ Problem	● Materials	● Evaluation (1, 2, 3)
○ Hypothesis	● Procedure	
● Prediction	● Evidence	

Problem

What mass of precipitate is produced by the complete reaction of 2.00 g of strontium nitrate in solution with an excess of aqueous copper(II) sulfate?

To perform this investigation, turn to page 305.

Applications of Stoichiometry

Having tested the stoichiometric method several times, you now have evidence that it can be accepted as valid and used with confidence to answer questions.

Calculating Percent Yield for Reactions

We can use stoichiometry to test experimental designs, technological skills, purity of chemicals (**Figure 7**), and the quantitative nature of a particular reaction. For each of these situations, we can evaluate the overall experiment by calculating a **percent yield**. This is the ratio of the actual or experimental quantity of product obtained (actual yield) to the maximum quantity of product (**theoretical yield** or predicted yield) obtained from a stoichiometry calculation:

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

Figure 7
Chemicals come in a wide variety of grades (purities). Some low-purity or technical grades may only be 80% to 90% pure, whereas high-purity or reagent grades may be better than 99.9% pure. The purity of a chemical can significantly affect experimental results when studying chemical reactions.

In laboratory work, many factors, called experimental uncertainties, can affect the percent yield of a chemical reaction. The only process you assume to be exact is stoichiometry itself. Some common sources of *experimental uncertainty* are the following:

- All measurements. Even assuming the experimenter makes careful measurements with the correct technique, inherent limitations in equipment always create some uncertainty. This also applies to initial values that are derived from measurements, such as molar masses and amount concentrations. For these values, the uncertainty is usually very small but it is not zero.
- The purity of the grade of chemical used (**Figure 7**). Where possible, consult the container label.
- Washing a precipitate. Very fine particles may be lost through the filter paper, or a very small amount of precipitate may be dissolved by repeated washings and lost that way.
- Any qualitative judgments that affect measurements. Estimation of colour or colour changes and estimation of reaction completion are two common examples.

It is not a simple matter to convert this list of experimental uncertainties into a numerical value, such as a percentage. *For school laboratories, investigations usually involve a total of all experimental uncertainties in the range of 5% to 10%.* This means that a percent yield as low as 90% could be considered quite acceptable for a particular experiment, depending on the equipment and chemicals used.

See Appendix B.2 for further tips on calculating and reporting percent yield.

Learning Tip

Scientists and technicians recognize and accept that there are many sources of experimental uncertainty, sometimes called sources of error. However, "human error" is not an acceptable category. If an experimenter makes a mistake, then the trial or experiment is repeated.

LAB EXERCISE 7.B

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2)
○ Hypothesis	○ Procedure	
○ Prediction	○ Evidence	

Testing a Chemical Process

Some technological problem solving involves quality control tests. These are physical and/or chemical tests performed during or at the end of a chemical process. The tests make sure that the process is working within parameters determined by the person in charge of quality control. In the Evaluation, evaluate the Design, list sources of experimental uncertainty, and then evaluate the Prediction.

Purpose

The purpose of this exercise is to perform a quality control test on a chemical process.

Problem

What is the mass of sodium silicate in a 25.0 mL sample of the solution used in a chemical process?

Prediction

If the process is operating as expected, the mass of sodium silicate in a 25.0 mL sample should always be between 6.40 g and 6.49 g.

Design

An excess quantity of iron(III) nitrate is added to the sample of sodium silicate. The resulting precipitate is separated by filtration. After the precipitate has dried, its mass is determined.

Evidence

mass of filter paper = 0.98 g
mass of dried filter paper plus precipitate = 9.45 g
The colour of the filtrate was yellow-orange.

Section 7.2 Questions

1. A balanced chemical equation includes simple coefficients in front of the chemical formulas.
(a) What do these coefficients represent?
(b) What is the term for the overall relationship of chemical amounts of all reactants and products?

2. List four assumptions about chemical reactions. Which two assumptions cannot be tested simply by observing a reaction?

3. In your own words, explain gravimetric stoichiometry.

4. How is a scientific concept such as stoichiometry tested? Provide a specific example.

5. For automobiles powered by hydrogen fuel cells to become successful, a source of hydrogen is required. Hydrogen can easily be produced by the electrolysis (simple decomposition) of water, but this process is very expensive.
(a) What perspective is being used to evaluate the production of hydrogen?
(b) Write the balanced chemical equation for the simple decomposition equation of water.
(c) Based on the coefficients in this chemical equation, if 100 g of hydrogen is produced, does this mean 50 g of oxygen will be formed? Justify your answer.

6. A chemical laboratory technician plans to react 3.50 g of lead(II) nitrate with excess potassium bromide in solution. Predict the mass of precipitate expected.

7. When calculating a percent yield for a reaction, where do the values for the actual yield and for the predicted yield come from?

8. In a chemical analysis, 3.00 g of silver nitrate in solution was reacted with excess sodium chromate to produce 2.81 g of filtered, dried precipitate.
(a) Using stoichiometry, predict the mass of precipitate expected in this reaction.
(b) Calculate the percent yield.

9. List four different sources of experimental uncertainty.

10. A solution made by dissolving 9.8 g of barium chloride is to be completely reacted with a second solution containing dissolved sodium sulfate.
(a) Predict the mass of precipitate expected.
(b) If 10.0 g of precipitate actually formed, calculate the percent yield.
(c) Does the percent yield result indicate the reaction went as expected?

11. Air-bag technology in automobiles has saved many lives. Research air-bag technology on the Internet and using other sources, such as newspapers and periodicals. Working in a group, prepare a presentation to explain how air-bags work. Your presentation should include the following:
- a list of the main chemicals and the main reactions
- an evaluation of air-bag technology
- an outline of the roles played by science and society in the development of this technology
- communication technology (in making your presentation)

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Extension

12. If you have access to the software, develop a spreadsheet that will predict the mass of a reagent required to yield various masses of product for a given reaction.

7.3 Gas Stoichiometry

Figure 1
Propane gas barbecues have become very popular. Charcoal barbecues are now banned in parts of California because they produce five times as much pollution (nitrogen oxides, hydrocarbons, and particulates) as gas barbecues.

Many chemical reactions involve gases. One common consumer example is the combustion of propane in a home gas barbecue (**Figure 1**). The reaction of chlorine in a water treatment plant is a commercial example. An important industrial application of a chemical reaction involving gases is the production of the fertilizer ammonia from nitrogen and hydrogen gases. These technological examples feature gases as either valuable products, such as ammonia, or as part of an essential process, such as water treatment.

Studies of chemical reactions involving gases (e.g., the law of combining volumes, Chapter 4) have helped scientists develop concepts about molecules and explanations for chemical reactions, such as the collision–reaction theory (Chapter 4). In both technological applications and scientific studies of gases, it is necessary to accurately calculate quantities of gaseous reactants and products.

The method of stoichiometry applies to all chemical reactions. This section extends stoichiometry to gases—**gas stoichiometry**—using gas volume, pressure and temperature, molar volume, and the ideal gas law.

Learning Tip

Symbols can be modified to clarify what they refer to by adding a subscript. For example, $n_{C_3H_8}$ represents the chemical amount of propane, and V_{O_2} represents the volume of oxygen.

SAMPLE problem 7.3

If 275 g of propane burns in a gas barbecue, what volume of oxygen measured at STP is required for the reaction?

First, write a balanced chemical equation to relate the amount of propane to the amount of oxygen. List the given and required values and the conversion factors for each chemical, just as you did in previous stoichiometry questions.

$C_3H_8(g)$	$+\ 5\,O_2(g)$	$\rightarrow$	$3\,CO_2(g) + 4\,H_2O(g)$
275 g	V (STP)		
44.11 g/mol	22.4 L/mol		

Since propane and oxygen are related by their mole ratio, you must convert the mass of propane to its chemical amount:

$$n_{C_3H_8} = 275\ \cancel{g} \times \frac{1\ \text{mol}}{44.11\ \cancel{g}}$$
$$= 6.23\ \text{mol}$$

The balanced equation indicates that 1 mol of propane reacts with 5 mol of oxygen. Use this mole ratio to calculate the amount of oxygen required, in moles. (This step is common to all stoichiometry calculations.)

$$n_{O_2} = 6.23\ \text{mol} \times \frac{5}{1}$$
$$= 31.2\ \text{mol}$$

Finally, convert the amount of oxygen to the required quantity, in this case, volume:

$$V_{O_2} = 31.2\ \cancel{\text{mol}} \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}}$$
$$= 698\ \text{L}$$

Note that the final step used the molar volume at STP as a conversion factor, in the same way that molar mass is used in gravimetric stoichiometry.

As in gravimetric stoichiometry, all steps may be combined as a single calculation:

$$V_{O_2} = 275\ \cancel{g\ C_3H_8} \times \frac{1\ \cancel{mol\ C_3H_8}}{44.11\ \cancel{g\ C_3H_8}} \times \frac{5\ \cancel{mol\ O_2}}{1\ \cancel{mol\ C_3H_8}} \times \frac{22.4\ L\ O_2}{1\ \cancel{mol\ O_2}}$$

$$= 698\ L\ O_2$$

COMMUNICATION example 1

Hydrogen gas is produced when sodium metal is added to water. What mass of sodium is necessary to produce 20.0 L of hydrogen at SATP?

Solution

$2\ Na(s)$ + $2\ H_2O(l)$	$\rightarrow$	$H_2(g)$ + $2\ NaOH(aq)$	
m		20.0 L	
22.99 g/mol		24.8 L/mol	

$$n_{H_2} = 20.0\ \cancel{L} \times \frac{1\ mol}{24.8\ \cancel{L}}$$

$$= 0.806\ mol$$

$$n_{Na} = 0.806\ mol \times \frac{2}{1}$$

$$= 1.61\ mol$$

$$m_{Na} = 1.61\ \cancel{mol} \times \frac{22.99\ g}{1\ \cancel{mol}}$$

$$= 37.1\ g$$

or

$$m_{Na} = 20.0\ \cancel{L\ H_2} \times \frac{1\ \cancel{mol\ H_2}}{24.8\ \cancel{L\ H_2}} \times \frac{2\ \cancel{mol\ Na}}{1\ \cancel{mol\ H_2}} \times \frac{22.99\ g\ Na}{1\ \cancel{mol\ Na}}$$

$$= 37.1\ g\ Na$$

According to gas stoichiometry, 37.1 g of sodium is needed to produce 20.0 L of hydrogen at SATP.

Learning Tip

Recall that when working with gas measurement, two sets of standard pressure and temperature conditions have been defined. These should be memorized. STP is a temperature of 0 °C and a pressure of 101.325 kPa (1 atm). SATP defines conditions closer to normal lab conditions: a pressure of 100 kPa and a temperature of 25 °C. These are exact values because they are definitions. Under conditions normal for laboratory work with gases, we assume the molar volume at STP is 22.4 L/mol and that the molar volume at SATP is 24.8 L/mol.

Note that the general steps of a stoichiometry calculation are the same for both solids and gases. Changes from mass to chemical amount or from volume to chemical amount, or vice versa, are done using the molar mass or the molar volume, respectively, of the substance. Although the molar mass depends on the chemical involved, the molar volume of a gas depends only on temperature and pressure. If the conditions are not standard (i.e., STP or SATP), then the ideal gas law ($PV = nRT$), rather than the molar volume, is used to find the amount or volume of a gas, as in the following example.

Learning Tip

Recall that a Celsius temperature scale is not useful for gas quantity calculations because it is not an absolute scale. In other words, it does not start from zero temperature. Gas law calculations use the (absolute) Kelvin scale of temperature. (Recall Section 4.1.) The exact conversion is as follows:

$T\ (K) = t\ (°C) + 273.15$

For most purposes, this conversion value may be rounded off to

$T\ (K) = t\ (°C) + 273$

which is accurate enough for all calculation questions in this textbook.

COMMUNICATION example 2

Ammonia, which is widely used as a fertilizer, is produced from the reaction of nitrogen and hydrogen. What volume of ammonia at 450 kPa pressure and 80 °C can be obtained from the complete reaction of 7.5 kg of hydrogen?

Solution

$N_2(g)$ +	$3\ H_2(g)$	$\rightarrow$	$2\ NH_3(g)$
	7.5 kg		V, 450 kPa, 353 K
	2.02 g/mol		8.314 kPa•L/(mol•K)

Learning Tip

The ideal gas law relationship is necessarily complex, because three variables, P, V, and T, must all be known to define a specific amount, n, of any gas. For gas stoichiometry questions that involve the constant R, it is often useful to write down the variation of the memorized formula $PV = nRT$ that is being applied.

Note in Communication Example 2 that care must be taken in any cancellation of unit values written with prefixes. The unit g cancels (from the known mass of H_2), but the prefix k does not.

$$n_{H_2} = 7.5\ \text{k}\cancel{\text{g}} \times \frac{1\ \text{mol}}{2.02\ \cancel{\text{g}}}$$
$$= 3.7\ \text{kmol}$$

$$n_{NH_3} = 3.7\ \text{kmol} \times \frac{2}{3}$$
$$= 2.5\ \text{kmol}$$

$$V_{NH_3} = \frac{nRT}{P}$$
$$= \frac{2.5\ \text{k}\cancel{\text{mol}} \times \dfrac{8.314\ \cancel{\text{kPa}} \bullet \text{L}}{1\ \cancel{\text{mol}} \bullet \cancel{\text{K}}} \times 353\ \cancel{\text{K}}}{450\ \cancel{\text{kPa}}}$$
$$= 16\ \text{kL}$$

or
$$V_{NH_3} = 7.5\ \text{k}\cancel{\text{g H}_2} \times \frac{1\ \cancel{\text{mol H}_2}}{2.02\ \cancel{\text{g H}_2}} \times \frac{2\ \cancel{\text{mol NH}_3}}{3\ \cancel{\text{mol H}_2}} \times \frac{8.314\ \cancel{\text{kPa NH}_3} \bullet \text{L NH}_3}{1\ \cancel{\text{mol NH}_3} \bullet \cancel{\text{K NH}_3}} \times \frac{353\ \cancel{\text{K NH}_3}}{450\ \cancel{\text{kPa NH}_3}}$$
$$= 16\ \text{kL NH}_3$$

According to gas stoichiometry, from the complete reaction of 7.5 kg of hydrogen one can obtain 16 kL of ammonia.

SUMMARY Gravimetric and Gas Stoichiometry

Step 1: Write a balanced chemical equation and list the measurements, unknown quantity symbol, and conversion factors for the measured and required substances.

Step 2: Convert the measured quantity to a chemical amount using the appropriate conversion factor.

Step 3: Calculate the chemical amount of the required substance using the mole ratio from the balanced equation.

Step 4: Convert the calculated chemical amount to the final quantity requested using the appropriate conversion factor.

Stoichiometry Calculations

measured quantity
solids/liquids m
gases V, T, P } → n

mole ratio ↓

required quantity
solids/liquids m
gases V, T, P } ← n

Practice

1. What volume of oxygen at STP is needed to completely burn 15 g of methanol in a fondue burner?
2. A Down's Cell is used in the industrial production of sodium from the decomposition of molten sodium chloride. A major advantage of this process compared with earlier technologies is the production of the valuable byproduct chlorine. What volume of chlorine gas at 30 °C and 95.7 kPa is produced, along with 105 kg of sodium metal, from the decomposition of sodium chloride?
3. Hydrogen gas is the fuel used in "pollution-free" vehicles in which hydrogen and oxygen gases react to produce water vapour and energy. Ballard Power Systems Inc. is a Canadian company pioneering the use of hydrogen fuel cells as power sources. Ballard heavy-duty fuel cells are currently being used to power a fleet of 30 Mercedes buses in trials in 10 European countries. What volume of oxygen at 40 °C and 1.50 atm is necessary to react with 300 L of hydrogen gas measured at the same conditions? (Recall the law of combining volumes.)

EXTENSION

Family Farming and Future Fuels

Ethanol created from agricultural waste can be used to produce hydrogen for fuel cells. This might give farmers a new source of income while providing a renewable and sustainable source of fuel. A research team at the University of Minnesota is investigating this process.

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INVESTIGATION 7.3 *Introduction*

Producing Hydrogen

There are several possible methods that can be used in the Design and Analysis. The suggested method is to predict the volume of gas at STP and, in your Analysis, convert the measured volume to STP conditions using the combined gas law.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
○ Hypothesis	○ Procedure	
● Prediction	● Evidence	

Purpose

The purpose of this investigation is to test the stoichiometric method applied to reactions that involve gases.

Problem

What is the volume at STP of hydrogen gas from the reaction of magnesium with excess hydrochloric acid?

Design

A known mass of magnesium ribbon reacts with excess hydrochloric acid. The temperature, pressure, and volume of the hydrogen gas produced are measured.

To perform this investigation, turn to page 305.

Case Study

Producing Hydrogen for Fuel Cells

Hydrogen fuel cells are promoted as being environmentally friendly because their only product is water vapour. This claim is true only if the hydrogen used in the cells is produced with the minimum of environmental impact. Hydrogen is found in many compounds that occur in nature, but the element is very difficult to isolate in a reliable, efficient, and economic way.

At present, the hydrocarbon molecules in fossil fuels, primarily natural gas, are the main source of hydrogen. The industrial process of reforming fossil fuels to make hydrogen is a method called *steam reforming* (**Figure 2**), in which vaporized fossil fuels react with steam at high pressures and temperatures in the presence of a nickel-based catalyst:

$$CH_4(g) + 2\,H_2O(g) \rightarrow CO_2(g) + 4\,H_2(g)$$

Figure 2
Most hydrogen is produced industrially by steam reforming.

The steam reforming process is well established, and it is currently the most economic way to produce hydrogen. Some disadvantages of this process are that it consumes energy and dwindling fossil fuels and produces carbon dioxide, the primary greenhouse gas.

Another process for isolating hydrogen gas uses electricity. Electrolysis produces hydrogen by using an electrical current to separate water into hydrogen and oxygen:

$$2\,H_2O(l) \rightarrow 2\,H_2(g) + O_2(g)$$

Unfortunately, the energy for electrolysis usually comes from burning fossil fuels, so again, carbon dioxide production is a problem, as well as using up the limited supply of fossil fuels. Producing hydrogen by electrolysis is much more environmentally friendly when solar or wind power is used as the energy source (**Figure 3**). The development of hydrogen generators powered by low-voltage sources from renewable-power technology is currently an area of active research.

Biomass, such as carbohydrate-rich agricultural rubbish and wastewater from food processing, is also being examined as a source of hydrogen (**Figure 4**). Hydrogen-producing bacteria, which occur naturally in soil, can be used to produce commercial quantities of hydrogen gas, for example,

$$C_6H_{12}O_6(aq) + 2\,H_2O(l) \rightarrow 2\,CH_3COOH(aq) + 2\,CO_2(g) + 4\,H_2(g)$$

Proponents of producing hydrogen from biomass point out that the carbon dioxide produced can be absorbed by planting more agricultural crops. Any industrial scale process for hydrogen generation must be as reliable, economic, and efficient as the reforming of fossil fuels. The success of the search could well determine whether hydrogen's promise as the clean fuel of the future will be fully realized.

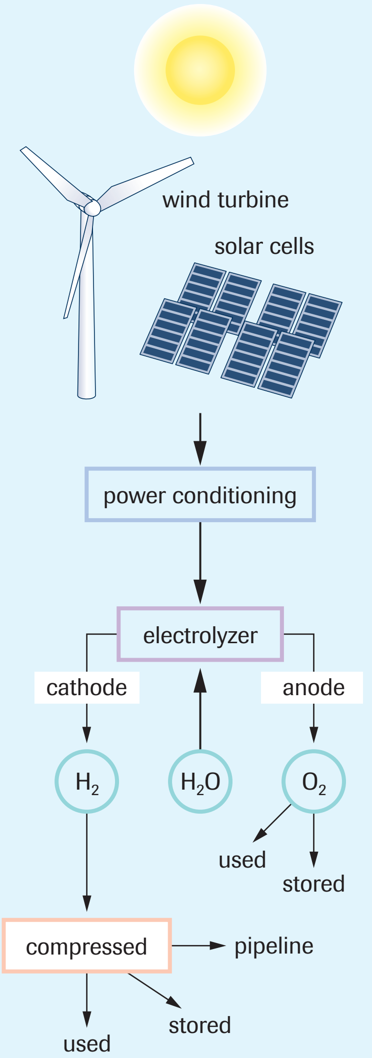

Figure 3
Solar-powered and wind-powered electrolysis is more environmentally friendly than electrolysis using fossil fuels.

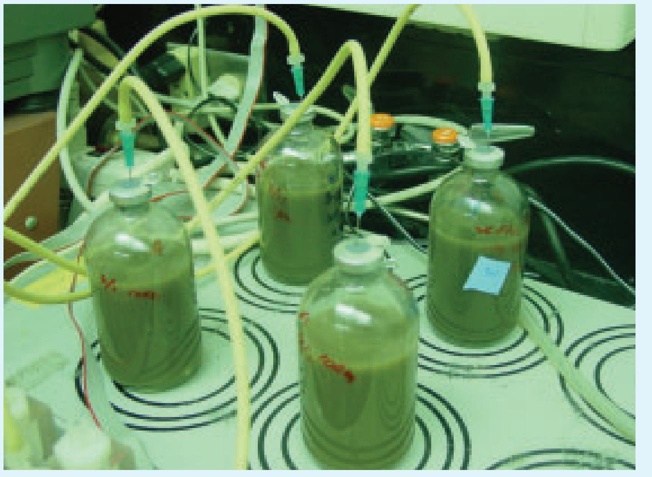

Figure 4
Hydrogen-producing bacteria can produce commercial quantities of hydrogen.

Case Study Questions

1. Based on the stoichiometry of the reactions given on the previous page, what are the ratios of hydrogen to starting material for each of the three processes? To evaluate hydrogen sources, is it better to compare chemical amount ratios or mass ratios?
2. For the production of hydrogen or any other industrial technological process, what are the three main criteria that are used to judge the process?
3. A source of hydrogen is an important issue if hydrogen fuel cells are going to become useful for automobiles. What perspectives are mentioned in the Case Study? What important perspective is missing?

Extension

4. Search the Internet and other sources for reports of recent research on a process that extracts hydrogen from water or biomass. Evaluate the potential for this technology from a variety of perspectives. Include an analysis of the risks and benefits expected, and discuss the long-term sustainability of the process in the presentation of your findings.

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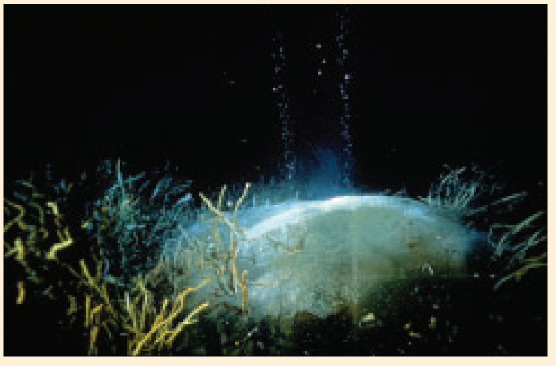
Section 7.3 Questions

1. How does gravimetric stoichiometry compare with gas stoichiometry? Identify the similarities and differences in the procedures.
2. The first recorded observation of hydrogen gas was made by the famous alchemist Paracelsus (1493–1541) when he added iron to sulfuric acid. Predict the volume of hydrogen gas at STP produced by adding 10 g of iron to an excess of sulfuric acid.
3. A typical Alberta home heated with natural gas (assume methane, $CH_4(g)$) consumes 2.00 ML of natural gas during the month of December. What volume of oxygen at SATP is required to burn 2.00 ML of methane measured at 0 °C and 120 kPa?
4. Ammonia reacts with sulfuric acid to form the important fertilizer ammonium sulfate. What mass of ammonium sulfate can be produced from 75.0 kL of ammonia at 10 °C and 110 kPa?

5. Methane hydrate, a possible energy resource, looks like ice but is an unusual substance with the approximate chemical formula $CH_4{\cdot}6H_2O(s)$. It occurs in permafrost regions and in large quantities on the ocean floor (**Figure 5(a)**). Current, rough estimates of the quantity of methane hydrate suggest that it is at least twice the total known reserves of coal, oil, and natural gas combined. Considerable research is now underway to find ways to tap this huge energy resource. If 1.0 kg of solid methane hydrate decomposes to methane gas and water, what volume of methane is produced at 20 °C and 95 kPa (**Figure 5(b)**)?

(a)



(b)

Figure 5

(a) In the frigid ocean depths extreme pressure forms this mound of methane hydrate. When the camera's light warms the mound, bubbles of methane can be seen dissociating from the ice.

(b) The methane escaping from this block of methane hydrate has been ignited, and it burns while the ice melts to water.

6. As recently as the early 20th century, pinches of sulfur were sometimes burned in sickrooms. The pungent choking fumes produced were supposed to be effective against the "evil humours" of the disease. In fact, the sulfur dioxide gas produced is toxic and extremely irritating to lung tissue, where it dissolves to form sulfurous acid. Even today, a surprising number of people still believe that medicines are more likely to be effective if they have unpleasant tastes or odours. What volume of $SO_2(g)$ at SATP will be produced from the burning of 1.0 g of sulfur?

Figure 6
Alberta has large supplies of sulfur as a byproduct of natural gas production. Sulfur is used to make sulfuric acid and ammonium sulfate.

7. Alberta's natural gas often has hydrogen sulfide gas, $H_2S(g)$, mixed with it (among other things) when it comes out of a well. Hydrogen sulfide is highly toxic and must be removed from the gas stream. In the second step of this removal process, hydrogen sulfide reacts with sulfur dioxide gas at high temperatures to produce water vapour and sulfur vapour. Upon cooling, the sulfur condenses to a solid, which is then stockpiled (**Figure 6**). If 1000 L of $H_2S(g)$ at SATP reacts in this way, what mass of solid sulfur would be formed?

8. In this test, aqueous hydrogen peroxide is decomposed to water and oxygen gas. Complete the Prediction and Evaluation (Part 2 only) sections of the following report.

Purpose
The purpose of this investigation is to test the stoichiometric method for gas reactions.

Problem
What volume of oxygen at room conditions can be obtained from the decomposition of 50.0 mL of 0.88 mol/L aqueous hydrogen peroxide?

Design
A measured volume of a hydrogen peroxide solution (3%, 0.88 mol/L) is decomposed using manganese dioxide as a catalyst. The oxygen produced is collected by water displacement, just like the hydrogen in Investigation 7.3.

Evidence
volume of 0.88 mol/L $H_2O_2(aq) = 50.0$ mL

volume of $O_2(g) = 556$ mL

temperature = 21 °C

atmospheric pressure = 94.6 kPa

9. Describe briefly one consumer, one industrial, and one laboratory application of gases that involve a chemical reaction that uses or produces gases. For each example, include a complete balanced chemical equation.

7.4 Solution Stoichiometry

You have already seen the usefulness of gravimetric stoichiometry and gas stoichiometry for both predictions and analyses. However, the majority of stoichiometric work in research and in industry involves solutions, particularly aqueous solutions. Solutions are easy to handle and transport, and reactions in solution are relatively easy to control.

Solution stoichiometry is the application of stoichiometric calculation principles to substances in solution. The general stoichiometric method remains the same. The major difference is that the amount concentration and volume of a solution are used as conversion factors to convert to or from the chemical amount of substance.

INVESTIGATION 7.4 *Introduction*

Analysis of Silver Nitrate (Demonstration)

It is more financially viable to recycle metals if they are in fairly concentrated solutions, so recycling companies will pay more for these solutions than for dilute solutions. How do companies find out how much silver, for example, is in a solution? Technicians carry out a reaction that involves removing all the silver from a known volume of the solution, drying it, and measuring its mass.

Purpose

The purpose of this investigation is to use the stoichiometric method to find an unknown amount concentration.

Problem

What is the amount concentration of silver nitrate in solution?

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ○ Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ● Evidence
- ● Analysis
- ● Evaluation (1, 2, 3)

Design

A precisely measured volume of aqueous silver nitrate solution, $AgNO_3$(aq), reacts completely with excess copper metal, Cu(s). The silver metal product, Ag(s), is separated by filtration and dried, and the mass of silver is measured to the precision of the balance. The amount concentration of the initial solution is calculated from the mass of product by the stoichiometric method.

To perform this investigation, turn to page 307.

Figure 1
Fertilizers can have a dramatic effect on plant growth. The plants on the left were fertilized with an ammonium hydrogen phosphate fertilizer.

SAMPLE problem 7.4

Solutions of ammonia and phosphoric acid are used to produce ammonium hydrogen phosphate fertilizer (**Figure 1**). What volume of 14.8 mol/L NH_3(aq) is needed for the ammonia to react completely with 1.00 kL of 12.9 mol/L H_3PO_4(aq) to produce fertilizer?

First, write a balanced chemical equation so that the stoichiometry can be established. Beneath the equation, list both the given and the required measurements and the conversion factors:

$2NH_3$(aq)	+	H_3PO_4(aq)	→	$(NH_4)_2HPO_4$(aq)
V		1.00 kL		
14.8 mol/L		12.9 mol/L		

Second, convert the information given for phosphoric acid to its chemical amount:

$$n_{H_3PO_4} = 1.00\ \text{k}\cancel{\text{L}} \times 12.9\ \frac{\text{mol}}{1\ \cancel{\text{L}}}$$

$$= 12.9\ \text{kmol}$$

Third, use the mole ratio to calculate the amount of the required substance, ammonia. According to the balanced chemical equation, 2 mol of ammonia reacts for every 1 mol of phosphoric acid:

$$n_{NH_3} = 12.9 \text{ kmol} \times \frac{2}{1}$$
$$= 25.8 \text{ kmol}$$

Fourth, convert the amount of ammonia to the quantity requested in the question. The amount concentration is used to convert the chemical amount to the solution volume:

$$V_{NH_3} = 25.8 \text{ k}\cancel{\text{mol}} \times \frac{1 \text{ L}}{14.8 \cancel{\text{mol}}}$$
$$= 1.74 \text{ kL}$$

As before, all steps may be combined as a single calculation:

$$V_{NH_3} = 1.00 \text{ k}\cancel{\text{L H}_3\text{PO}_4} \times \frac{12.9 \cancel{\text{mol H}_3\text{PO}_4}}{1 \cancel{\text{L H}_3\text{PO}_4}} \times \frac{2 \cancel{\text{mol NH}_3}}{1 \cancel{\text{mol H}_3\text{PO}_4}} \times \frac{1 \text{ L NH}_3}{14.8 \cancel{\text{mol NH}_3}}$$
$$= 1.74 \text{ kL NH}_3$$

COMMUNICATION *example*

A technician determines the amount concentration, *c*, of a sulfuric acid solution. In the experiment, a 10.00 mL sample of sulfuric acid reacts completely with 15.9 mL of 0.150 mol/L potassium hydroxide solution. Calculate the amount concentration of the sulfuric acid.

Solution

$H_2SO_4(aq)$ +	$2\,KOH(aq) \rightarrow 2\,H_2O(l) + K_2SO_4(aq)$
10.00 mL	15.9 mL
c	0.150 mol/L

$$n_{KOH} = 15.9 \text{ m}\cancel{\text{L}} \times \frac{0.150 \text{ mol}}{1 \cancel{\text{L}}}$$
$$= 2.39 \text{ mmol}$$

$$n_{H_2SO_4} = 2.39 \text{ mmol} \times \frac{1}{2}$$
$$= 1.19 \text{ mmol}$$

$$[H_2SO_4(aq)] = \frac{1.19 \cancel{\text{m}}\text{mol}}{10.00 \cancel{\text{m}}\text{L}}$$
$$= 0.119 \text{ mol/L}$$

or

$$[H_2SO_4(aq)] = 15.9 \cancel{\text{m}} \cancel{\text{L KOH}} \times \frac{0.150 \cancel{\text{mol KOH}}}{1 \cancel{\text{L KOH}}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \cancel{\text{mol KOH}}} \times \frac{1}{10.00 \cancel{\text{m}}\text{L H}_2\text{SO}_4}$$
$$= \frac{0.119 \text{ mol H}_2\text{SO}_4}{1 \text{ L H}_2\text{SO}_4}$$
$$= 0.119 \text{ mol/L H}_2\text{SO}_4$$

According to the stoichiometric method for solutions, the amount concentration of the sulfuric acid is 0.119 mol/L.

DID YOU KNOW ?

Plant Food

Nitrogen is the primary nutrient for plant growth. It promotes protein formation in crops and is a major component of chlorophyll, which helps promote healthy growth, producing high yields.

CAREER CONNECTION

Soil Scientist

Soil scientists use chemistry to investigate the composition of soil and how it behaves. They study how people and industry affect soil, and they create and monitor plans to remediate contaminated soils. Soil scientists also work closely with the agricultural and forestry industries to study and predict interactions of soil with other organisms.

Soil scientists are critical to understanding our environment. Find out whether this could be the career for you.

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SUMMARY Gravimetric, Gas, and Solution Stoichiometry

Stoichiometry Calculations

measured quantity
solids/liquids *m*
gases *V, T, P* → *n*
solutions *c, V*

↓ mole ratio

required quantity
solids/liquids *m*
gases *V, T, P* ← *n*
solutions *c, V*

Step 1: Write a balanced chemical equation, and list the quantities and conversion factors for the given substance and the one to be calculated.

Step 2: Convert the given measurement to its chemical amount using the appropriate conversion factor.

Step 3: Calculate the amount of the other substance using the mole ratio from the balanced equation.

Step 4: Convert the calculated amount to the final quantity requested using the appropriate conversion factor.

Practice

1. Ammonium sulfate fertilizer is manufactured by having sulfuric acid react with ammonia. In a laboratory study of this process, 50.0 mL of sulfuric acid reacts with 24.4 mL of a 2.20 mol/L ammonia solution to produce the ammonium sulfate solution. From this evidence, calculate the amount concentration of the sulfuric acid at this stage in the process.
2. Slaked lime can be added to an aluminium sulfate solution in a water treatment plant to clarify the water. Fine particles in the water stick to the precipitate produced. Calculate the volume of 0.0250 mol/L calcium hydroxide solution required to react completely with 25.0 mL of 0.125 mol/L aluminium sulfate solution.
3. A chemistry teacher wants 75.0 mL of 0.200 mol/L iron(III) chloride solution to react completely with an excess quantity of 0.250 mol/L sodium carbonate solution. What is the minimum volume of sodium carbonate solution needed?

LAB EXERCISE 7.C

Testing Solution Stoichiometry

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ● Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ○ Evidence
- ● Analysis
- ● Evaluation (1, 2, 3)

You have already tested the stoichiometric method for gravimetric and gas stoichiometry, but the testing of a scientific concept is never finished. Scientists keep looking for new experimental designs and new ways of testing a scientific concept. When completing the investigation report, pay particular attention to the evaluation of the Design.

Purpose

The purpose of this exercise is to test the stoichiometric method using solutions.

Problem

What mass of precipitate is produced by the reaction of 20.0 mL of 0.210 mol/L sodium sulfide with an excess quantity of aluminium nitrate solution?

Design

The two solutions provided react with each other, and the resulting precipitate is separated by filtration and dried. The mass of the dried precipitate is determined.

Evidence

A yellow precipitate resembling aluminium sulfide was formed.
mass of filter paper = 0.97 g
mass of dried filter paper plus precipitate = 1.17 g
A few additional drops of the sodium sulfide solution added to the filtrate produced a precipitate. *Hint*: What compound do you expect to be present in the filtrate solution?

LAB EXERCISE 7.D

Determining a Solution Concentration

Once a scientific concept has passed several tests, it can be used in industry. In your career as an industrial technician, you need to determine the amount concentration of a silver nitrate solution that, due to its cost, is being recycled.

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ○ Prediction
- ● Design
- ○ Materials
- ○ Procedure
- ○ Evidence
- ● Analysis
- ○ Evaluation

Purpose

The purpose of this exercise is to use the stoichiometric method with solutions.

Problem

What is the amount concentration of silver nitrate in the solution to be recycled?

Evidence

A white precipitate was formed in the reaction with aqueous sodium sulfate.

volume of silver nitrate solution = 100 mL

mass of filter paper = 1.27 g

mass of dried filter paper plus precipitate = 6.74 g

Section 7.4 Questions

1. Some antacid products contain aluminium hydroxide to neutralize excess stomach acid. Determine the volume of 0.10 mol/L stomach acid (assumed to be HCl(aq)) that can be neutralized by 912 mg of aluminium hydroxide in an antacid tablet.
2. Sulfuric acid is produced on a large scale from readily available raw materials. One step in the industrial production of sulfuric acid is the reaction of sulfur trioxide with water. Calculate the amount concentration of sulfuric acid produced by the reaction of 10.0 Mg of sulfur trioxide with an excess quantity of water to produce 7.00 kL of acid.
3. Analysis shows that 9.44 mL of 50.6 mmol/L KOH(aq) is needed to completely react with 10.00 mL of water from an acidic lake. Determine the amount concentration of acid in the lake water, assuming that the acid is sulfuric acid.
4. Silver nitrate solution is used by electroplating businesses to replate silver tableware for their customers (**Figure 2**). To test the purity of the solution, a technician adds 10.00 mL of 0.500 mol/L silver nitrate to an excess quantity of 0.480 mol/L NaOH solution. From the reaction, 0.612 g of precipitate is obtained.
   (a) State a specific diagnostic test that could be done to verify that an excess had been added.
   (b) Calculate the predicted yield of precipitate.
   (c) What is the percent yield? What does this tell you about the purity of the solution?
5. Some commercial hydrochloric acid mixed with water produces 20.0 L of a 1.20 mol/L solution to be used to remove rust from car parts in a wrecking yard. What mass of rust can be reacted before the acid is used up? Assume solid Fe_2O_3 as a formula for rust.
6. Design an experiment to determine the amount concentration of a sodium sulfate solution. Include the Problem, Design, and Materials.

Figure 2
A small-scale electroplating business

Extension

7. In the late 1800s, two chemical processes, LeBlanc and Solvay, competed as methods for producing soda ash. The Solvay process clearly won and remains a major chemical industry today.
   (a) Why is soda ash important? List some uses.
   (b) What are the two major raw materials and two final products of the Solvay process?
   (c) Is the Solvay process a consumer-, commercial-, or industrial-scale technology?
   (d) Why was the Solvay process so successful? Working in a team, evaluate this technology compared to the LeBlanc process. Include a variety of perspectives.

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INVESTIGATION 7.1

Report Checklist		
○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
○ Hypothesis	● Procedure	
● Prediction	● Evidence	

Decomposing Malachite

Copper(II) hydroxide carbonate, commonly called basic copper carbonate and also known as malachite, is a double salt with the chemical formula $Cu(OH)_2 \cdot CuCO_3(s)$ (**Figure 1**). This double salt decomposes completely when heated to 200 °C, forming copper(II) oxide, carbon dioxide, and water vapour. Complete the Prediction using the balanced chemical equation. Include safety and disposal steps in your Procedure. Organize the data and create the graph using suitable software.

Purpose

The purpose of this investigation is to test the assumption that a chemical reaction is stoichiometric.

Problem

How is the chemical amount of copper(II) oxide produced related to the chemical amount of malachite reacted in the decomposition of malachite?

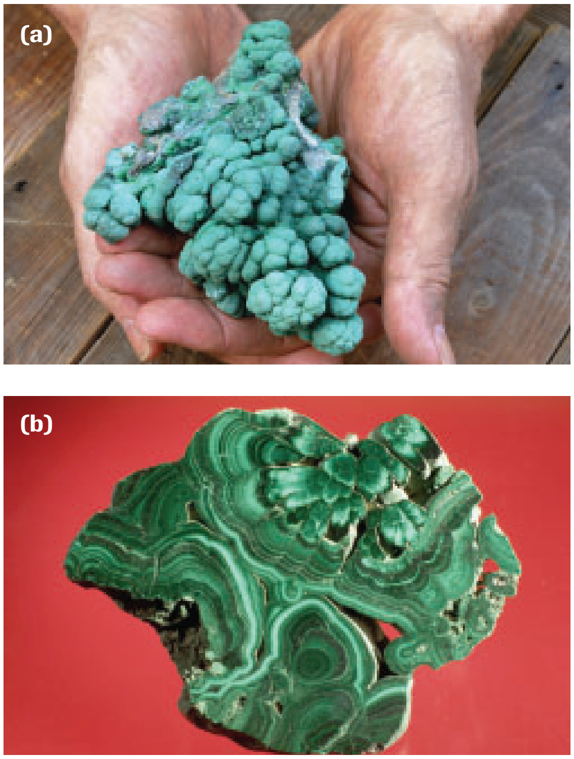

Figure 1
(a) The green mineral malachite is an important copper ore.
(b) When polished, it is also used as a semi-precious stone in jewellery. When prepared as a pure chemical substance, it is usually ground to a light green powder.

Design

A known mass of malachite (manipulated variable) is heated strongly until the colour changes completely from green to black (**Figure 2**). The mass of black copper(II) oxide (responding variable) is determined. The results from several laboratory groups are combined in a graph to determine the ratio of chemical amounts.

Materials

lab apron
eye protection
porcelain dish (or crucible and clay triangle)
small ring stand
hot plate
glass stirring rod
sample of malachite (1 g to 3 g)
centigram balance
laboratory scoop or plastic spoon

Malachite is toxic.
Do not touch the surface of the hot plate.

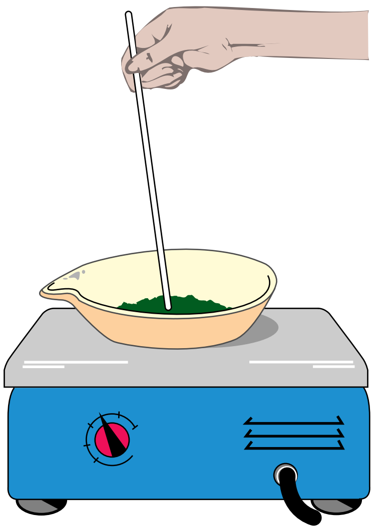

Figure 2
Use the glass stirring rod to break up lumps of powdered malachite and to mix the contents of the dish while they are being heated. Large lumps may decompose on the outside but not on the inside.

INVESTIGATION 7.2

Gravimetric Stoichiometry

In this investigation, you will use gravimetric stoichiometry to investigate the reaction of strontium nitrate with excess copper(II) sulfate in an aqueous solution. Use 2.00 g of strontium nitrate and about 3.5 g of copper(II) sulfate–water (1/5), initially dissolving each chemical in about 75 mL of water. Be sure to include safety and disposal instructions in your Procedure. Refer to Appendix C.4 and the Nelson Web site for guidance on various lab techniques.

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Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ● Prediction
- ● Design
- ● Materials
- ● Procedure
- ● Evidence
- ● Analysis
- ● Evaluation (1, 2, 3)

Purpose
The purpose of this investigation is to test the stoichiometric method.

Problem
What mass of precipitate is produced by the complete reaction of 2.00 g of strontium nitrate in solution with an excess of aqueous copper(II) sulfate?

Strontium nitrate is moderately toxic; there is risk of fire when it is in contact with organic chemicals, and it may explode when bumped or heated. Copper(II) sulfate is a strong irritant and is toxic if ingested.

INVESTIGATION 7.3

Producing Hydrogen

There are several possible methods that can be used in the Design and Analysis. The suggested method is to predict the volume of gas at STP and, in your Analysis, convert the measured volume to STP conditions using the combined gas law.

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ○ Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ● Evidence
- ● Analysis
- ● Evaluation (1, 2, 3)

Purpose
The purpose of this investigation is to test the stoichiometric method applied to reactions that involve gases.

Problem
What is the volume at STP of hydrogen gas from the reaction of magnesium with excess hydrochloric acid?

Design
A known mass of magnesium ribbon reacts with excess hydrochloric acid. The temperature, pressure, and volume of the hydrogen gas produced are measured.

Eye protection, a lab apron, and disposable gloves must be worn.

Hydrochloric acid in 6 mol/L concentration is very corrosive. If acid is splashed into your eyes, rinse them immediately with water for 15 to 20 min. Acid splashed onto the skin should be rinsed immediately with plenty of water. Notify your teacher immediately. If acid is splashed onto your clothes, neutralize with baking soda, then wash thoroughly with plenty of water.

Rinse your hands well after step 8 in case any dilute acid got on your skin.

Hydrogen gas, produced in the reaction of hydrochloric acid and magnesium, is flammable. Ensure that there is adequate ventilation and that there are no open flames in the classroom.

Materials
lab apron
eye protection
disposable plastic gloves
magnesium ribbon, 60 mm to 70 mm
centigram or analytical balance
piece of fine copper wire, 100 mm to 150 mm
100 mL graduated cylinder
15 mL hydrochloric acid (6 mol/L)
250 mL beaker
water
large beaker (600 mL or 1000 mL)

INVESTIGATION 7.3 *continued*

two-hole stopper to fit cylinder
thermometer or temperature probe
barometer

Procedure

1. Measure and record the mass of the strip of magnesium.
2. Fold the magnesium ribbon to make a small compact bundle that can be held by a copper cage (**Figure 3**).

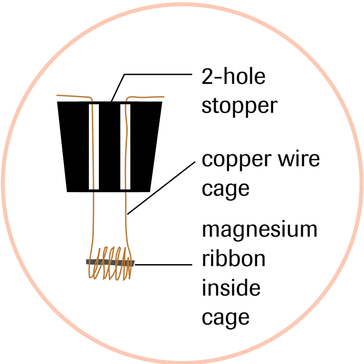

Figure 3
The magnesium should be small enough to fit into a copper cage (steps 2 and 3). Fasten the copper wire handle to the stopper (step 7).

3. Wrap the fine copper wire all around the magnesium, making a cage to hold it, but leaving 30 mm to 50 mm at each end of the wire free for a handle.
4. Carefully pour 10 mL to 15 mL of the hydrochloric acid into the graduated cylinder.
5. Slowly fill the graduated cylinder to the brim with water from a beaker. As you fill the cylinder, pour slowly down the side of the cylinder to minimize mixing of the water with the acid at the bottom. In this way, the liquid at the top of the cylinder is relatively pure water and the acid remains at the bottom.
6. Half-fill the large beaker with water.
7. Bend the copper wire handle through the holes in the stopper so that the cage holding the magnesium is positioned about 10 mm below the bottom of the stopper (**Figure 3**).
8. Insert the stopper into the graduated cylinder; the liquid in the cylinder will overflow a little. Cover the holes in the stopper with your finger. Working quickly, invert the cylinder, and immediately lower it into the large beaker so that the stopper is below the surface of the water before you remove your finger from the stopper holes (**Figure 4**).

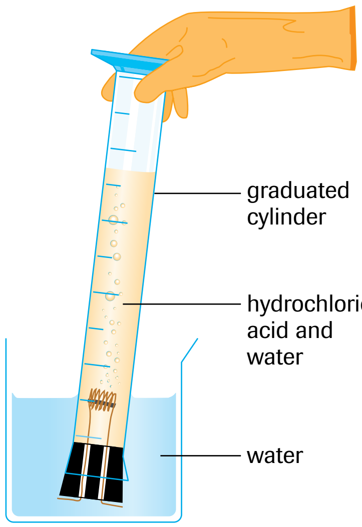

Figure 4
While holding the cylinder so it does not tip, rest it on the bottom of the beaker. The acid, which is denser than water, will flow down toward the stopper and react with the magnesium. The hydrogen produced should remain trapped in the graduated cylinder.

9. Observe the reaction, then wait about 5 min after the bubbling stops to allow the contents of the graduated cylinder to reach room temperature.
10. Raise or lower the graduated cylinder so that the level of liquid inside the beaker is the same as the level of liquid inside the graduated cylinder. (This equalizes the gas pressure in the cylinder with the pressure of the air in the room.)
11. Measure and record the volume of gas in the graduated cylinder.
12. Record the laboratory (ambient) temperature and pressure.
13. The liquids in this investigation may be poured down the sink, but rinse the sink with lots of water.

INVESTIGATION 7.4

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
○ Hypothesis	○ Procedure	
○ Prediction	● Evidence	

Analysis of Silver Nitrate (Demonstration)

It is more financially viable to recycle metals if they are in fairly concentrated solutions, so recycling companies will pay more for these solutions than for more dilute solutions. How do companies find out how much silver, for example, is in a solution? Technicians carry out a reaction that involves removing all the silver from a known volume of the solution, drying it, and measuring its mass.

Purpose

The purpose of this investigation is to use the stoichiometric method to find an unknown amount concentration.

Problem

What is the amount concentration of silver nitrate in solution?

Design

A precisely measured volume of aqueous silver nitrate solution, $AgNO_3(aq)$, is completely reacted with excess copper metal, Cu(s). The silver metal product, Ag(s), is separated by filtration and dried, and the mass of silver measured to the precision of the balance. The amount concentration of the initial solution is calculated from the mass of product by the stoichiometric method.

Materials

lab apron
eye protection
>100 mL $AgNO_3(aq)$ of unknown amount concentration
centigram balance
#16–#20 gauge solid (not braided) copper wire
fine steel wool
wash bottle of pure water
wash bottle of pure acetone, $CH_3COCH_3(l)$
filtration apparatus
filter paper
250 mL beaker, with watch glass to fit
400 mL waste beaker for acetone
400 mL waste beaker for filtrate
100 mL graduated cylinder
stirring rod

Acetone is volatile and flammable. Use only in a well-ventilated area. Keep away from any source of flame or sparks.

Procedure

Many of the skills and techniques required for this investigation are described in Appendix C.

Day 1

1. Using a graduated cylinder, measure 100 mL of silver nitrate solution and pour it into a 250 mL beaker.
2. Clean about 30 cm of solid copper wire with fine steel wool, and form about 20 cm of it into a coil with a 10 cm handle, so the coiled section will be submerged when placed in the silver nitrate solution.
3. Record any immediate evidence of chemical reaction, cover with a watch glass, and set aside.

Day 2

4. Check for completeness of reaction. If the coil is intact, with unreacted (excess) copper remaining, the reaction is complete; proceed to step 6. If all the copper has reacted, proceed to step 5.
5. Add another coil of copper wire, cover with a watch glass, and set aside until the next day.
6. Remove the wire coil. Shake the coil to ensure that all silver crystals remain in the beaker.
7. Measure and record the mass of a piece of filter paper.
8. Filter the beaker contents to separate the solid silver from the filtrate. (This technique is described in Appendix C.4, and demonstrated in a video on the Nelson Web site.)

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9. Do the final three washes of the solid silver and filter paper with acetone from a wash bottle. Catch the rinsing acetone in a waste beaker.
10. Place the unfolded filter paper and contents on a paper towel to dry for a few minutes.
11. Measure and record the mass of the dry silver plus filter paper.
12. Dispose of solids in the garbage and the aqueous solutions (not acetone) down the drain with plenty of water. Transfer the acetone to a flammables disposal container.

Chapter 7 SUMMARY

Outcomes

Knowledge

- identify limitations and assumptions about chemical reactions (7.1)
- write balanced ionic and net ionic equations, including identification of spectator ions, for reactions taking place in aqueous solutions (7.1)
- recognize limiting and excess reagents in chemical reactions (7.1, 7.2, 7.3, 7.4)
- calculate quantities of reactants and/or products involved in chemical reactions using gravimetric, solution, or gas stoichiometry (7.2, 7.3, 7.4)
- define predicted (theoretical) and experimental (actual) yields, and explain the discrepancy between them (7.2, 7.3)
- identify sources of experimental uncertainty in experiments (7.2, 7.3, 7.4)

STS

- state that a goal of technology is to solve practical problems (7.2, 7.3, 7.4)
- recognize that technological problem solving may incorporate knowledge from various fields (7.2, 7.3)
- classify and evaluate technologies (7.2, 7.3, 7.4)
- explain how the appropriateness and the risks and benefits of technologies need to be assessed for each potential application from a variety of perspectives, including sustainability (7.3)

Skills

- initiating and planning: plan and predict states, products, and theoretical yields for chemical reactions (7.2); describe procedures for safe handling, storing, and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information (7.2, 7.4)
- performing and recording: translate word equations for chemical reactions into chemical equations, including states of matter for the products and reactants (7.2); balance chemical equations for chemical reactions, using lowest whole-number coefficients (7.2)
- analyzing and interpreting: interpret stoichiometric ratios from chemical reaction equations (7.2, 7.3, 7.4); perform calculations to determine theoretical yields and actual yields, percent yield, and error (7.2); use appropriate SI notation, fundamental and derived units, and significant digits when performing stoichiometry calculations (7.2, 7.3, 7.4)
- communication and teamwork: work collaboratively in addressing problems and applying the skills and conventions of science in communicating information and ideas and in assessing results (7.2)

Key Terms

7.1
quantitative reaction
stoichiometric reaction
net ionic equation
spectator ion
limiting reagent
excess reagent

7.2
stoichiometry
theoretical yield
gravimetric stoichiometry
percent yield

7.3
gas stoichiometry

7.4
solution stoichiometry

Key Equations

$$\text{percent yield} = \frac{\text{actual yield}}{\text{predicted yield}} \times 100 \qquad (7.2)$$

MAKE a summary

1. Expand the margin summary graphic on page 302 to clearly show the following:
   (a) the three systems used for initially calculating the amount of a measured substance and how many separate measurements are required for each
   (b) the six possible required final quantities from a stoichiometric calculation and the system used to calculate each of them
2. 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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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 7
- 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 this chapter.

EXTENSION

CBC radio ONE QUIRKS & QUARKS

Touchy-Feely Robots
One of the obstacles stopping robots from performing delicate manipulations is their lack of touch sensitivity. Researchers are working on a new material, incorporating nanoparticles and electrodes, that will give robots the sense of touch.

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Chapter 7 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.

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DO NOT WRITE IN THIS TEXTBOOK.

Part 1

1. A main goal of technology is to
   A. advance science
   B. identify problems
   C. explain natural processes
   D. solve practical problems

2. In the reaction of aqueous solutions of sodium sulfide and zinc nitrate in a chemical analysis, the spectator ions are
   A. sodium and nitrate ions
   B. sulfide and zinc ions
   C. sodium and zinc ions
   D. sulfide and nitrate ions

3. In which sections of an investigation report are stoichiometric calculations most likely to be found?
   A. Problem and/or Procedure
   B. Prediction and/or Analysis
   C. Purpose and/or Materials
   D. Hypothesis and/or Evaluation

4. (NR) The four general steps of any stoichiometry calculation are given in the following numbered list.
   1. converting a chemical amount to another quantity
   2. writing a balanced equation and listing information
   3. converting another quantity (or more than one) to a chemical amount
   4. determining one chemical amount from another chemical amount

   List the order in which these steps occur:
   ____, ____, ____, and ____.

Use this information to answer questions 5 to 9.

The mineral malachite is mined for use as a copper ore (**Figure 1**). After malachite has been roasted (decomposed by heat), the next step in the production of copper metal is a single replacement reaction. Copper(II) oxide reacts with hot carbon to produce copper metal and carbon dioxide. Assume that a 1.00 kg sample of pure copper(II) oxide is reacted.

Figure 1
If copper ore is close to the surface, it is recovered using huge digging and transport machinery in an open pit mine.

5. (NR) The chemical amount of the copper(II) oxide to be reacted is __________ mol.

6. (NR) The mass of carbon that will be required to completely react with all the copper(II) oxide is __________ g.

7. This reaction situation suggests the use of an excess reagent. The substance that should be deliberately supplied in excess quantity is
   A. $CO_2(g)$
   B. $CuO(s)$
   C. $C(s)$
   D. $Cu(s)$

8. (NR) The mass of copper that should be formed by completely reacting all the copper(II) oxide is __________ g.

9. The carbon dioxide produced is vented to the atmosphere. What volume would this amount of carbon dioxide occupy at SATP?
   A. 156 L
   B. 77.9 L
   C. 39.0 L
   D. 24.8 L

Use this information to answer questions 10 to 13.

The balanced equation for the combustion of sulfur is $S_8(s) + 8\,O_2(g) \rightarrow 8\,SO_2(g)$. Assume a sample of pure sulfur is burned in air, which is about 20% oxygen.

10. The balanced equation provides you with clear and direct information about
    A. the temperature and pressure at which the reaction will be spontaneous
    B. the likelihood that the reaction will be quantitative (complete)
    C. the initial rate of reaction and the time required for the reaction to finish
    D. the ratio of chemical amounts of reactants and products

11. At a glance, without any calculation, it is possible to confidently say that
    A. the mass of oxygen that reacts will be eight times the mass of sulfur that reacts
    B. the chemical amount of sulfur dioxide formed is greater than the amount of sulfur reacting
    C. oxygen will be the limiting reagent
    D. sulfur dioxide gas will form at SATP conditions

12. (NR) The mass of toxic sulfur dioxide gas produced by quantitative reaction of 4.00 g of sulfur is __________ g.

13. If this reaction were to be done inside a sealed container, with the only change in conditions being the use of 100% pure oxygen, the reaction should happen much faster because the rate of collisions between oxygen molecules and sulfur molecules should become much greater. What volume of oxygen, measured at SATP, would be required to burn each 1.00 kg of sulfur under these conditions?
    A. 773 L
    B. 544 L
    C. 224 L
    D. 24.8 L

Part 2

14. List the four basic assumptions made for chemical reactions when doing stoichiometric calculations.

15. What is meant by the percent yield of a reaction, for which a stoichiometric calculation was used to predict the quantity of a product?

16. List the common sources of experimental uncertainty that may account for some of the difference between predicted and experimental quantities.

17. **Explain** why, in all stoichiometric calculations, you always have to convert to or convert from chemical amounts.

18. DE Technology has always been a part of any society, even going back to the Stone Age. As knowledge, societal needs, and problems increase, more sophisticated technology develops.
(a) Technologies may be classified according to their scale and use. What are three contexts used to classify technology?
(b) For a particular technology, what are the main criteria used to judge the product or process?
(c) List at least five perspectives that may be used when evaluating technologies.

19. DE A sodium phosphate solution is used to test tap water for the presence of calcium ions (**Figure 2**). A sample of tap water reacts with sodium phosphate solution to produce a precipitate.
(a) Write the net ionic equation for the reaction.
(b) **Identify** the spectator ion(s).
(c) Based on the given design, **identify** the limiting and excess reagents.
(d) **Identify** a possible, significant flaw in the design of this experiment.

Figure 2
The precipitate, formed when clear, colourless sodium phosphate solution is added to tap water, indicates that the tap water contains dissolved water-hardening “impurities.”

20. DE When heated, baking soda (**Figure 3**) decomposes into solid sodium carbonate, carbon dioxide, and water vapour.
(a) If 2.4 mol of baking soda decomposes, what chemical amount of each of the products is formed?
(b) What mass of solid product will remain after complete decomposition of a 1.00 kg box of baking soda?
(c) Suggest a reason why baking soda can be used as a fire extinguisher.

Figure 3
Baking soda is pure sodium bicarbonate. An open box should always be in any kitchen because it is an excellent extinguisher for small cooking fires.

21. A convenient source of oxygen in a laboratory is the decomposition of aqueous hydrogen peroxide to produce water and oxygen. What volume of 0.88 mol/L hydrogen peroxide solution (**Figure 4**) is required to produce 500 mL of oxygen at SATP?

Figure 4
When some coloured liquid soap is added, hydrogen peroxide decomposes rapidly to produce water and oxygen. The soap foams as it traps bubbles of the oxygen gas being formed.

Use this information to answer questions 22 to 25.

In a chemical reaction done to test the stoichiometric method, 3.00 g of silver nitrate in aqueous solution reacts with a large excess of sodium chromate in solution to produce 2.81 g of dry precipitate.

22. **Determine** the predicted mass of precipitate.

23. (a) What is the percent yield?
(b) **Predict** some reasons that might account for the difference between the predicted and the actual yield.

24. Write the net ionic equation for this analysis.

25. **Identify** the spectator ions.

26. In plants, the process called photosynthesis involves a reaction to produce glucose and oxygen from carbon dioxide and water (**Figure 5**). This endothermic reaction is powered by light energy from the sun and is catalyzed by chlorophyll:

$$6\ CO_2(g) + 6\ H_2O(l) + \text{energy} \rightarrow C_6H_{12}O_6(aq) + 6\ O_2(g)$$

(a) **Predict** the mass of carbon dioxide consumed when a plant makes 10.0 g of glucose.
(b) **Predict** the mass of oxygen produced when a plant makes 10.0 g of glucose.

Figure 5
The production of carbohydrates by plants is the fundamental source of energy for almost all living things on Earth.

Use this information to answer questions 27 to 32.

Gasohol is a general term that refers to automobile fuel that has 10% ethanol, $C_2H_5OH(l)$, blended with unleaded gasoline (**Figure 6**). Using this fuel reduces some noxious exhaust emissions. Assume that 1.00 kg of ethanol reacts completely in a car engine.

Figure 6
Blending gasoline with ethanol not only produces a fuel that burns more cleanly, and also helps conserve petroleum resources.

27. Write the balanced equation for the complete combustion of ethanol.

28. **Predict** the mass of oxygen is required?

29. **Predict** the mass of carbon dioxide produced.

30. **Predict** the mass of water produced.

31. Show that your calculated answers to the previous questions agree with the law of conservation of mass.

32. **Illustrate**, using examples involving automobile transportation, polluting emissions, and catalytic converters, how science and technology have both intended and unintended consequences and also how science often leads technology, and technology often leads science.

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33. A metal refinery that uses a hydrometallurgical (aqueous solution) process to produce pure metals uses a stoichiometric procedure to determine the cobalt(II) ion amount concentration in the process solution. Complete the Analysis of the investigation report.

Purpose
The purpose of this investigation is to use stoichiometric calculations to determine an unknown solution amount concentration.

Problem
What is the amount concentration of cobalt(II) sulfate in a 100.0 mL sample of process solution?

Design
Solid sodium carbonate is added to the hot aqueous sample and dissolved, with stirring. When adding more sodium carbonate causes no more precipitate to form, the precipitate is allowed to settle and is then filtered. The mass of the dried precipitate is determined.

Evidence
A red crystalline precipitate was formed in the reaction.

volume of cobalt(II) sulfate solution	100.0 mL
mass of filter paper	1.04 g
mass of dry precipitate plus paper	8.98 g

34. Make a list of theories, laws, generalizations, and rules that you must know in order to be able to solve a stoichiometry problem.

35. Chemical technicians in water treatment plants perform several routine reactions daily on a very large scale. Research and report on the use of stoichiometry for ensuring the quality of a municipal water supply. Your report should include
- descriptions of two or three water-quality tests involving stoichiometry
- graphics
- properly referenced data

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