Chapter

9

The Gas State

The photograph in **Figure 1** is a dramatic depiction of how a gas can save a human life. In a car crash, an air bag, especially in combination with a seat belt, can protect the driver from serious injury. Upon collision, sensors in the steering column and in the bumper initiate the decomposition of sodium azide into sodium metal and nitrogen gas. This reaction is extremely fast: Nitrogen gas is produced and expands into the bag in less than 0.04 s. After cushioning the impact, the air bag gradually deflates as the nitrogen gas escapes through the permeable bag. Instead of taking a trip to the hospital, the driver takes a trip to the automobile body shop to have the air bag mechanism recharged and the triggering devices reset.

Air bags are not the only use of gases in the operation of automobiles: Tires and shock absorbers are inflated with pressurized air to provide a safe and comfortable ride. Air enters through the car's vents and is cooled by the air conditioner to keep us comfortable on hot summer days or is heated by the car engine to keep us warm in winter. Inside the combustion cylinders of the engine, a gasoline and oxygen explosion produces a large amount of gas at high temperature, which moves a piston. This is an example of converting chemical energy into motion. Finally, the gases emitted by automobile exhausts, such as carbon oxides and nitrogen oxides, diffuse into the atmosphere as pollutants.

As you can see, gases play an important role in both technology and our natural environment. In this chapter, you will learn more about the properties and uses of gases.

In this chapter, you will be able to

- explain different states of matter in terms of the forces among atoms, molecules, and ions;
- use the kinetic molecular theory to describe and explain the behaviour of gases;
- determine through experimentation the algebraic and graphical relationships among the pressure, volume, and temperature of an ideal gas;
- describe the mathematical relationships among the pressure, volume, temperature, and amount of an ideal gas;
- solve quantitative problems involving laws that describe the properties and behaviour of gases;
- use the terms standard temperature and pressure (including STP and SATP), absolute or Kelvin temperature, and ideal gas;
- convert between various units of pressure and between Celsius and Kelvin temperatures;
- describe various natural events and technological products and processes associated with gases;
- identify technological uses and safety concerns of compressed gases;
- identify the components of the atmosphere and describe Canadian initiatives to improve air quality.

Reflect on your Learning

1. Since many gases are invisible, how do you think we can study them?
2. All around us, we see examples of all three states of matter. Why are some substances solid, liquid, or gas? How is this explained by the forces between the molecules?
3. Weather reports often refer to low- and high-pressure systems. What does pressure of a gas mean?
4. What determines the quality of the air in our atmosphere?

Throughout this chapter, note any changes in your ideas as you learn new concepts and develop your skills.

Try This **Activity**

Creating and Testing Ideas About Gases

Coming up with ideas to explain experiments with gases requires some imagination, because most gases are invisible. Can you figure out what is happening to the aluminum cans and water vapour (a gas in our atmosphere) in this activity?

Materials: water, 5 aluminum pop cans, 5 large beakers or containers, hot plate, beaker tongs, ice cubes, eye protection

Care is required handling hot items. Steam can scald skin. Switch off hot plate immediately after use.

- Place about 20 mL of water in an empty aluminum pop can.
- Heat the can on a hot plate until steam rises steadily out of the top for a couple of minutes.
- Fill a large beaker to near the top with cold water.
- Using the tongs, lift the can and move it quickly to the beaker of cold water.
- Invert the can, and dip the top rim of the can just under the surface of the water.
- Record your observations.

  (a) Create a Hypothesis for what happens.

- Repeat the Procedure without placing any water in the can. (Heat the can for a few minutes.)

  (b) What happens now? Does this support or refute your Hypothesis?

  (c) Using your original or revised Hypothesis, predict the results if you repeat the Procedure inverting the steaming can into ice water and warm water.

  (d) Try each of these, then judge your Prediction and Hypothesis.

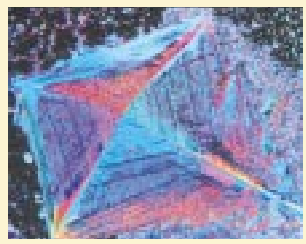
- Recycle the cans.

Figure 1
Air bags are a good example of how knowledge of gas reactions and gas properties can be used in life-saving technology.

9.1 States of Matter

The empirical properties (**Table 1**) of the three states of matter—solid, liquid, gas—provide important clues about the nature and structure of matter. At the same temperature some substances are solids, some are liquids, and some are gases. This suggests that the forces between the particles that make up the substance vary in strength. Unlike liquids and gases, solids maintain their shape and volume (**Figure 1**). This evidence suggests that the forces between the particles in the solid are strong; in fact, these forces are the strongest of the three states. At room temperature, all ionic compounds and all metals (except mercury) are solid. Some molecular substances, including both molecular elements and compounds, can also be in the solid state. The ionic bonds in ionic compounds, intermolecular forces in some molecular substances, covalent bonds in covalent crystals, and metallic bonds in metallic substances all provide strong attractions that bond the particles tightly together. Strong bonds are also believed to prevent solids from flowing easily, since this would require particles to be able to slip past one another. Strong bonds would also suggest that there are few empty spaces between the particles, and this would explain why solids are difficult to compress.

Table 1: Empirical Properties of States of Matter

State	Properties	Example
solid	• solids have definite shape and volume • are virtually incompressible • do not flow easily	 **Figure 1** A crystal of table salt
liquid	• liquids assume the shape of the container but have a definite volume • are virtually incompressible • flow readily	 **Figure 2** Coloured water
gas	• gases assume the shape and volume of the container • are highly compressible • flow readily	 **Figure 3** Bicycle air pump

At room temperature, most liquids are molecular compounds. This evidence plus their generally low boiling points suggest that the intermolecular forces (dipole–dipole, London, and hydrogen bonding) are not as strong as ionic, covalent, or metallic bonds. However, in these liquids, the intermolecular forces must be sufficiently strong to hold the molecules closely together but not locked in place, allowing them to move past one another. Therefore, the molecules in a liquid can spread out to take the shape of the container while keeping their volume constant (**Figure 2**). Because gases have no definite shape or volume (**Figure 3**), there appears to be an absence of forces between the molecules in a gas.

Atomic theory predicts that noble gases are composed of monatomic molecules (e.g., $Ne_{(g)}$). Since these molecules are nonpolar, intermolecular force theory suggests that any attraction between noble gas molecules must be explained by London (dispersion) forces. The larger the noble gas molecule, the larger the London forces, due to the larger number of electrons per molecule. Similarly, the intermolecular attractions between diatomic molecules of elements, for example, $H_{2(g)}$ and $Cl_{2(g)}$, can be explained by London forces. The van der Waals attraction between diatomic molecules of compounds, for example, $HCl_{(g)}$ and $CO_{(g)}$, is explained as being a combination of London and dipole–dipole forces. In general, because of large intermolecular distances in a gas, hydrogen bonding is not possible between gas molecules.

An attempt to explain the states of matter based only on the strengths of forces fails when considering how solids can be changed to liquids and liquids changed to gases by increasing the temperature. The reverse changes in state also occur when the temperature decreases. The strengths of bonds between molecules cannot be the complete explanation for the different states. To more completely understand the states of matter, we must also consider the motion of the molecules.

The Kinetic Molecular Theory

How would you explain why a drop of food colouring added to a glass of cold water slowly spreads out, or diffuses, throughout the water? How would you explain why the amount of water in an open container slowly decreases as some of the water evaporates? For the first question, scientists would say that the molecules of food colouring and the molecules of water are moving and colliding with each other, and this causes them to mix. The answer to the second question is also molecular motion: Some of the water molecules in the open container obtain sufficient energy from collisions to escape from the liquid. The idea of molecular motion that is used to explain these observations has led to the **kinetic molecular theory,** which has become a cornerstone of modern science.

The fundamental idea of the kinetic molecular theory is that solids, liquids, and gases are composed of particles that are continually moving and colliding with other particles. These particles may be atoms, ions, or molecules. As they move about, they collide with each other and with objects in their path. Very tiny objects, such as pollen grains or specks of smoke, are buffeted by these particles in air and move erratically, as shown in **Figure 4**.

There are three types of motion that any particle can exhibit: translational (straight-line), rotational (spinning), and vibrational (back-and-forth motion of atoms within the molecule). Which type of motion predominates depends on the freedom of movement of the particles; this, in turn, depends on the strengths of the forces between the particles. If particles are restricted to mainly vibrational motion, as in solids, then particles stay together in a relatively ordered state

DID YOU KNOW ?

Atmospheres

The Earth has an atmosphere, Saturn's moon Titan has an atmosphere, but the Moon does not. Why? This can be explained, at least partially, by kinetic molecular theory and gravitational attraction. The higher the temperature, the faster the particles in an atmosphere would move. The faster they move, the more likely they are to escape the pull of the body's gravity. (Molecules, like rockets, must reach escape velocity to leave an atmosphere.) Cool bodies are more likely to have an atmosphere, as are massive bodies, which pull more strongly on atmospheric particles. The Earth and Moon are the same distance from the Sun. Without an atmosphere their temperatures would be roughly the same. However, Earth is much more massive than the Moon, and so can retain an atmosphere. Saturn's moon Titan, although its mass is less than that of Earth, has a denser atmosphere. Temperatures are much lower so far from the Sun.

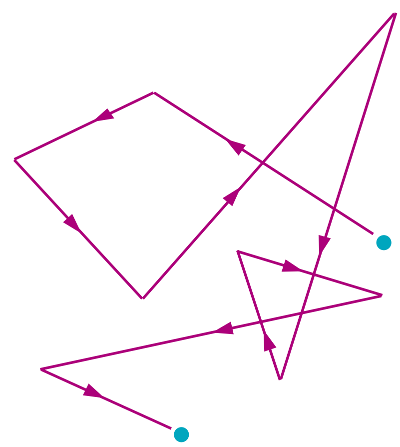
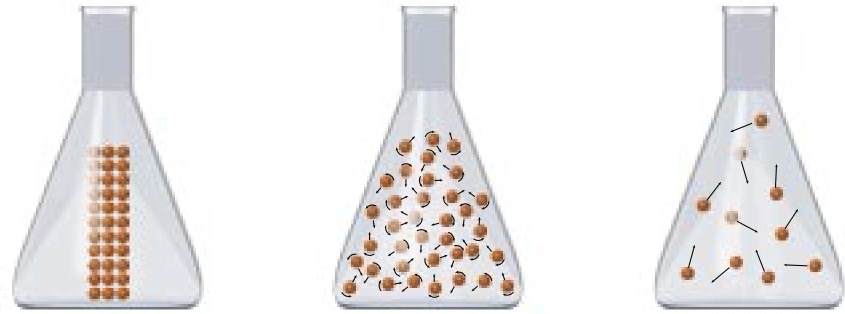


Figure 4
Observation of microscopic particles such as pollen grains or specks of smoke shows a continuous, random motion known as Brownian motion, named for Scottish scientist Robert Brown (1773–1858), who first described it. Scientists' interpretations of this evidence led to the formation of the kinetic molecular theory.

kinetic molecular theory: the idea that all substances contain particles that are in constant, random motion

Figure 5
According to the kinetic molecular theory, the motion of molecules is different in solids, liquids, and gases. Particles in solids have primarily vibrational motion; particles in liquids have vibrational, rotational, and translational motion; and the most important form of motion in gases is translational.



DID YOU KNOW ?

Expansion upon Heating

Most solids, liquids, and gases expand (their density decreases) when the temperature increases. This is explained as an increase in the average kinetic energy of the particles, which carries the particles farther apart before the intermolecular forces can pull them back. Of course, the increase in density when ice melts is explained differently. Do you know the explanation?

(**Figure 5**). Liquids with some of each type of motion remain largely together but in a more jumbled, less orderly state. Gas molecules rotate and vibrate but their translational (straight-line) motion is most significant. This produces random collisions and the most disordered state with no organization.

Any moving object has energy called kinetic energy. A moving car, bird, and molecule all have kinetic energy. The faster the motion of an object, the greater its kinetic energy. Because the molecules of a substance are always colliding, at any instant some molecules are moving faster than others. Therefore, in a large group of molecules, there will be a range of kinetic energies from very low to very high values. The temperature of a substance is a measure of the average kinetic energy of its particles. If a substance is heated, its temperature rises until a change of state occurs. As the temperature increases, the average kinetic energy of the particles increases and, on average, each particle moves faster and has more kinetic energy. When the average kinetic energy exceeds a particular (threshold) energy, a substance changes state; for example, melts or boils. The high-energy particles escape, and with a constant energy input from the outside of the system, the temperature remains constant until the change of state is complete.

Explaining the Gas State

Theoretical chemists have created theories to explain the empirical properties of the gas state. They explain that for solids and liquids, the intermolecular forces hold the particles together. These forces of attraction are opposed by forces of repulsion when the particles get too close together. We say that the particles collide. If the intermolecular forces are relatively strong and the kinetic energy is relatively low, then a condensed state is favoured. The properties of gases (**Table 1**) suggest that the intermolecular forces are virtually nonexistent. Consider a closed bottle of perfume. Initially, the perfume molecules are contained within the bottle. If the bottle is opened in a closed room, some of the perfume molecules that are continually escaping from the solution through evaporation can now leave the open bottle. The perfume gas molecules will slowly diffuse throughout their new container—the room. Opening the door allows the perfume molecules to eventually occupy the whole building. And opening the front door of the building means that the perfume gas molecules can diffuse into the atmosphere, a very large container. There does not appear to be a limit to the diffusion of gas molecules as long as they do not react with other substances. Based on this evidence and other properties of gases (e.g., lack of a definite shape), we will assume that there are no attractive forces between gas molecules and that the molecules move in straight lines independently of each other. This is a very good approximation when starting the study of gases.

Activity 9.1.1

Molecular Motion

Models in science are used to describe abstract ideas and make them easier to understand. Some models are mathematical (e.g., mathematical models of motion describing falling bodies), some are mechanical (e.g., model cars and trains), and some are physical (e.g., physical models of the biological cell). It is important to remember that all models have limitations. The model is good if the descriptions, explanations, and predictions it gives for natural phenomena outweigh its limitations.

In this activity, you will use a molecular motion demonstrator to better visualize and understand the models of solids, liquids, and gases.

Figure 6
A molecular motion demonstrator is a mechanical device that uses small spheres to represent molecules and agitation to simulate molecular motion.

Materials

molecular motion demonstrator (**Figure 6**)
watch glass
plastic spheres
overhead projector

Procedure

1. Set the molecular motion demonstrator on the overhead projector and place the watch glass inside the metal square. Add some plastic spheres to the watch glass.
2. Turn on the power to the demonstrator and set it to low speed. Observe and record the shape and volume of the whole sample and the motion of the individual spheres.
3. Slowly increase the speed until the spheres start to leave the main group. Observe and record the characteristics of the sample and the individual spheres.
4. Remove the watch glass and place the plastic spheres inside the main compartment. Slowly increase the speed until it is at the maximum setting. Record your observations.

Analysis

(a) How does the molecular model demonstrator simulate the attractive forces between particles of a solid or liquid?

(b) Based on your observations, describe the degree of order or organization of the particles in each model of the three states of matter.

Synthesis

(c) Evaluate the ability of this model to represent the molecular motion within solids, liquids, and gases.

(d) How could you use this model to show diffusion within a liquid or gas? the compressibility of gases?

DID YOU KNOW ?

Distances Between Molecules

In a molecular solid, the distance between molecules is about the same size as the molecules themselves; in a liquid it is slightly greater; and in a gas the distance between molecules is about 20 to 30 times the size of the molecules. To picture yourself as a particle in a solid, imagine yourself seated in a regular classroom. For a liquid, picture a school dance. For a gas, imagine yourself and three friends skating randomly at the Air Canada Centre.

Practice

Understanding Concepts

1. Arrange the three states of matter in order of increasing strengths of forces between the particles. Explain your order using the evidence of shapes and volumes.
2. Which state of matter has the highest degree of order and which has the lowest? How is this related to the forces between the particles in each state?
3. The approach of starting with a single observation, such as "gases are compressible," and asking a series of "why" questions illustrates the limits of our current theories.
   (a) Why are gases compressible? Answer this question in terms of the model of a gas.
   (b) Use your answer to (a) to ask and answer a further "why" question.
   (c) How far can you extend this series of questions? Try this and stop when you cannot answer a "why" question.

Making Connections

4. Hydraulic devices, such as the brake system of a car, have a piston at one end that pushes on a liquid connected by a hose to a piston at the other end of the closed system.
   (a) What property of a liquid allows this system to work?
   (b) Why is it dangerous if the liquid (brake fluid) leaks out and is replaced by an air bubble?

Reflecting

5. Models are useful to help visualize and understand abstract ideas. Different kinds of models, such as models that use physical objects, diagrams, and mathematical equations, often appeal to different people. What kind of model works best for you? If presented with one kind of model, do you sometimes try to switch to another kind?

Section 9.1 Questions

Understanding Concepts

1. (a) List the four classes of chemical substances that are or can be solids at room temperature.
   (b) Identify the types of bonds for each class.
   (c) Explain briefly why these substances are in the solid state.
2. Solids and liquids are often referred to as condensed states.
   (a) What empirical property is the same for both these states? Explain briefly.
   (b) What properties are different? Explain the differences in terms of forces and motion.
3. For substances that are gases at room temperature,
   (a) what interpretation can be made about the forces between the molecules?
   (b) what is the predominant type of motion for gas molecules?
4. Using appropriate theoretical concepts, explain each of the empirical properties of gases listed in **Table 1**, page 418.

Making Connections

5. What properties of gases make the air bag in an automobile useful as a safety device? For each property chosen, show how this applies to air bags.

9.2 Gas Laws

Gases have always been important to us—we need them to breathe after all—but as our society has advanced technologically, the importance of gases has been expanding. We use gases in our daily lives—natural gas as fuel, gases as refrigerants, anesthetic gases for surgery—and we generate gases for special investigation, for example we create artificial atmospheres for deep-sea diving and for the exploration of outer space. It is not surprising then that the study of gases has a long history in chemistry. Many experimental properties of gases were studied long before the development of our modern understanding of the composition and molecular motion of substances. In fact, some important ideas such as atomic theory, kinetic molecular theory, and the mole concept were made possible to a large extent by the large body of empirical knowledge about gases. Let's now look more closely at the empirical properties of gases.

Try This Activity

A Simulation of Gas Properties

Suppose five nitrogen gas cylinders are assembled using the conditions listed in **Table 1**. Each cylinder contains the same mass of nitrogen gas.

Table 1: Comparison of Nitrogen Gas Cylinders

Cylinder number	Volume (L)	Temperature (°C)
1	1.0	800
2	2.0	200
3	2.0	300
4	4.0	200
5	4.0	800

(a) What is the order of gas cylinders from most likely to least likely to explode? Write your Prediction and provide your reasoning for the order you choose.

(b) If you were designing and testing cylinders for the safe and efficient transportation of various gases, which variables would you need to consider?

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pressure: force per unit area

DID YOU KNOW ?

Standard Pressure

Another way to think about standard pressure is as the equivalent to the weight of one kilogram resting on every square centimetre of your body.

atmospheric pressure: the force per unit area exerted by air on all objects

Pressure and Volume: Boyle's Law

We live at the bottom of an ocean of air. That air has many different properties that can be altered experimentally in a laboratory, including temperature and **pressure** (properties familiar to us from weather reports), volume, and amount of gas. In any controlled experiment, the plan is to manipulate one variable and observe its effect on another variable while keeping all other properties constant. We begin our study of gases by looking at the relationship between pressure and volume at a constant temperature and amount of gas.

Earth's gravity exerts a downward force on you, and you, in turn, exert an equal force on the ground. However, the force you exert can be distributed over a large or a small area. The area is large when you lie down, and small when you stand on the tips of your toes. The greater the area, the lower the pressure. For example, when you wear snowshoes, the force is distributed over the surface area of the snowshoe, so you exert less pressure on the ground directly below your feet than you would if you were wearing regular shoes. This allows you to walk over snow instead of sinking into it. Pressure of a gas is also force per unit area, but in this case the force is exerted by the moving molecules as they collide with objects in their path, particularly the walls of a container.

Scientists have agreed, internationally, on units, symbols, and standard values for pressure. The SI unit for pressure, pascal (Pa), represents a force of 1 N (newton) on an area of 1 m^2; 1 Pa = 1 N/m^2. Atmospheric pressure and the pressure of many gases are often more conveniently measured in kilopascals (kPa); 1 kPa = 1000 Pa = 1 kN/m^2 (exactly).

At sea level, average **atmospheric pressure** is about 101 kPa. Scientists used this value as a basis to define one standard atmosphere (1 atm), or *standard pressure,* as exactly 101.325 kPa. For convenience, *standard ambient pressure* has been more recently defined as exactly 100 kPa.

For many years, standard conditions for work with gases were a temperature of 0°C and a pressure of 1 atm (101.325 kPa); these conditions are known as standard temperature and pressure (STP). However, 0°C is not a convenient temperature, because laboratory temperatures are not close to 0°C. Scientists have since agreed to use another set of standard conditions, not only for gases but also for reporting the properties of other substances. The new standard is called standard ambient temperature and pressure (SATP), defined as 25°C and 100 kPa. The new standard is much closer to laboratory conditions.

Since the empirical properties of gases were measured long before the development of SI, pressure of gas has been expressed in a bewildering variety of units over the years. In 1643, Evangelista Torricelli (1608–1647), following up on a suggestion from Galileo, accidentally invented a way of measuring atmospheric pressure. He was investigating Aristotle's notion that nature abhors a vacuum. His experimental design involved inverting a glass tube filled with mercury and placing it into a tub also containing mercury (**Figure 1**). Noticing that the mercury level changed from day to day, he realized that his device, which came to be called a mercury barometer, was a means of measuring atmospheric pressure. In Torricelli's honour, standard pressure was at one time defined as 760 torr, or 760 mm Hg. (Mercury vapour is toxic; in modern mercury barometers, a thin film of water or oil is added to prevent the evaporation of mercury from the open reservoir.)

Many areas of study that employ gases, such as medicine and meteorology, and several technological applications, such as deep-sea diving, still use non-SI units (**Table 2**). Using the definitions in **Table 2**, it is possible to easily convert between SI and non-SI units.

Table 2: SI and Non-SI Units of Gas Pressure

Unit name	Unit symbol	Definition/Conversion
pascal	Pa	1 Pa = 1 N/m^2
atmosphere	atm	1 atm = 101.325 kPa (exactly)
millimetres of mercury	mm Hg	760 mm Hg = 1 atm = 101.325 kPa
torr	torr	1 torr = 1 mm Hg

Sample Problem 1

Convert standard ambient pressure, defined as 100 kPa, to the corresponding values in atmospheres and millimetres of mercury.

Solution

$$100\ \cancel{\text{kPa}} \times \frac{1\ \text{atm}}{101.325\ \cancel{\text{kPa}}} = 0.987\ \text{atm}$$

$$100\ \cancel{\text{kPa}} \times \frac{760\ \text{mm Hg}}{101.325\ \cancel{\text{kPa}}} = 750\ \text{mm Hg}$$

Practice

Understanding Concepts

1. Define STP and SATP.
2. Copy and complete **Table 3**. Show your work using appropriate conversion factors.

Table 3: Converting Pressure Units

	Pressure (kPa)	Pressure (atm)	Pressure (mm Hg)
(a)	0.50		
(b)	96.5		
(c)			825
(d)		2.50	

3. What are the advantages of having only one unit for pressure?

Making Connections

4. When using a medicine dropper or a meat baster, you squeeze the rubber bulb and insert the end of the tube into a liquid. Why does the liquid rise inside the dropper or baster when you release the bulb?

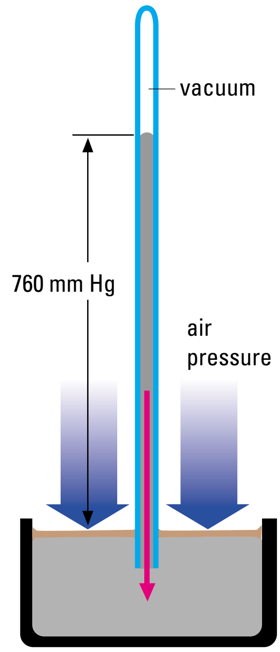

Figure 1
When a tube filled with mercury is inverted, the weight of the column of mercury pulls it toward Earth. However, the weight of air directly above the open dish pushes down on the surface of the mercury and prevents all of the mercury from falling out of the tube. The two opposing forces—weight of mercury and weight of air—balance each other when the height of mercury is about 760 mm. If the tube of mercury is longer than 760 mm, the mercury drops, leaving a vacuum above. Why is mercury used in most barometers, and not other liquids such as water, which is plentiful and nontoxic? The answer is density: Mercury is much denser than water. The weight of air in the atmosphere will support a column of water about 10 m high, which might be difficult to fit in a room!

Answers

2. (a) 0.0049 atm, 3.8 mm Hg
   (b) 0.952 atm, 724 mm Hg
   (c) 110 kPa, 1.09 atm
   (d) 253 kPa, 1.90×10^3 mm Hg

INQUIRY SKILLS

- ○ Questioning
- ○ Hypothesizing
- ○ Predicting
- ● Planning
- ● Conducting
- ● Recording
- ● Analyzing
- ● Evaluating
- ● Communicating

Investigation 9.2.1

Pressure and Volume of a Gas

The purpose of this investigation is to determine the general relationship between the pressure and volume of a gas. Complete the **Experimental Design, Analysis,** and **Evaluation** sections of the lab report.

Question

What effect does increasing the pressure have on the volume of a gas?

Experimental Design

(a) Using the Procedure and **Figure 2**, write a brief plan to summarize this experiment.
(b) Identify the independent, dependent, and two controlled variables.
(c) Design a table to record your observations.

Materials

Boyle's law apparatus or 35-mL plastic syringe
large rubber stopper
cork borer
5 textbooks or equal masses (1 kg)
utility stand
buret clamp
mass balance

Procedure

1. Pull out the syringe plunger so that 30 mL of air is inside the cylinder.
2. If a syringe cap is not provided, bore a small hole deep enough in the rubber stopper so that the tip of the syringe is inside the stopper. This should be a tight fit. Make sure the tip of the syringe does not leak.
3. Hold the syringe barrel vertical and measure the initial volume. Clamp the syringe on a retort stand.
4. While holding the syringe securely, carefully place one textbook or mass on the end of the plunger (**Figure 2**). (Your partner should balance the mass and be prepared to catch it if it starts to tilt.) Record the mass and new volume of air.
5. Repeat step 4 for a total of 4 or 5 books or masses.
6. If time permits, repeat steps 3 to 5 for an additional one or two trials.

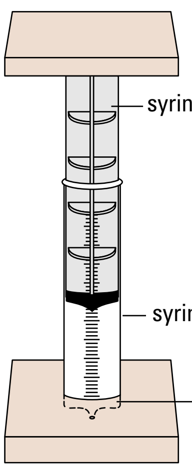

Figure 2
Setup of Boyle's law apparatus

Analysis

(d) Plot a graph of gas volume (or average volume from trials) versus mass added and draw a best-fit line.
(e) How does changing the mass on the syringe plunger affect the pressure on the air inside the syringe?
(f) According to the evidence you have collected, what effect does increasing pressure have on the volume of a gas?

Evaluation

(g) What are some sources of experimental error or uncertainty in this experiment? In your judgment, are these major or minor problems?
(h) How does your graph provide some indication of experimental errors or uncertainties in this experiment?
(i) Suggest some improvements that might raise the quality of the Evidence. Be as specific as possible.

The Relationship Between Pressure and Volume

Analysis of the evidence produced in an investigation similar to Investigation 9.2.1 suggests an inverse variation between the pressure and the volume of a gas; that is, as the pressure increases, the volume decreases (**Figure 3**). Using the evidence given in SI units in **Table 4**, you can see that when the pressure is doubled (100 kPa to 200 kPa), the volume is halved (3.00 L to 1.50 L). If the pressure is tripled, the volume is reduced to one-third. Check the other values to see similar results.

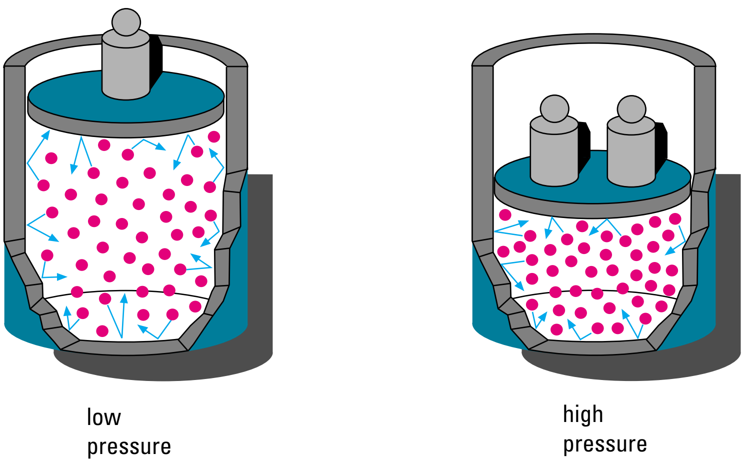

Figure 3
As the pressure on a gas increases, the volume of the gas decreases.

Table 4: Pressure and Volume of Gas Samples

Pressure (kPa)	Volume (L)	*pv* (kPa·L)
100	3.00	300
200	1.52	304
300	1.01	303
400	0.74	296
500	0.60	300

If p_1 and v_1 represent the initial conditions, the other values of pressure and volume from **Table 4** may be stated as follows:

$$(p_1, v_1) \quad (2p_1, \tfrac{1}{2}v_1) \quad (3p_1, \tfrac{1}{3}v_1) \quad (4p_1, \tfrac{1}{4}v_1) \quad (5p_1, \tfrac{1}{5}v_1)$$

For all the conditions listed above, the product of the pressure and volume is equal to p_1v_1. Mathematically, the relationship is represented as $pv = k$, where k is a constant. This simple relationship was first determined by Robert Boyle in 1662 (**Figure 4,** page 428). **Boyle's law** states that *as the pressure on a gas increases, the volume of the gas decreases proportionally, provided that the temperature and amount of gas remain constant.* In other words, the volume of a gas is inversely proportional to the pressure of the gas, providing that the temperature and amount of gas are held constant. Boyle's law can be conveniently written comparing any two sets of pressure and volume measurements:

$$p_1v_1 = p_2v_2 \quad \textit{(Boyle's law)}$$

This can also be expressed as a calculation of a new pressure inversely related to the volumes ratio:

$$p_2 = \frac{p_1v_1}{v_2}$$

Boyle's law: as the pressure on a gas increases, the volume of the gas decreases proportionally, provided that the temperature and amount of gas remain constant; the volume and pressure of a gas are inversely proportional

Figure 4
Anglo-Irish chemist Robert Boyle (1627–1691) determined the effect of pressure on the volume of a gas in quantitative terms. He was a founding member of the Royal Society of London and is reported to have coined its anti-Aristotelian motto: "Nothing by Authority." In the early 1660s, Boyle worked with Robert Hooke, the able inventor, who helped him construct an air pump. Using this necessary technology, Boyle demonstrated the physical characteristics of air and the necessary role of air in combustion, respiration, and the transmission of sound. In 1661, he reported to the Royal Society on the relationship now known as Boyle's law. Boyle became so famous that foreign academics wouldn't consider a trip to England complete until they had met him. He was elected president of the Royal Society in 1680, but declined the honour.

Sample Problem 2

A 2.0-L party balloon at 98 kPa is taken to the top of a mountain where the pressure is 75 kPa. Assume the temperature is the same. What is the new volume of the balloon?

Solution

$v_1 = 2.0$ L
$p_1 = 98$ kPa
$p_2 = 75$ kPa
$v_2 = ?$

$$p_1v_1 = p_2v_2$$

$$v_2 = \frac{p_1v_1}{p_2}$$

$$= \frac{98\ \cancel{\text{kPa}} \times 2.0\ \text{L}}{75\ \cancel{\text{kPa}}}$$

$$v_2 = 2.6\ \text{L}$$

or $v_{\text{balloon}} = 2.0\ \text{L} \times \frac{98\ \cancel{\text{kPa}}}{75\ \cancel{\text{kPa}}} = 2.6\ \text{L}$

The new volume of the balloon is 2.6 L.

Practice

Understanding Concepts

5. Define atmospheric pressure.
6. A bicycle pump contains 0.650 L of air at 101 kPa. If the pump is closed, what pressure is required to change the volume to 0.250 L?
7. A weather balloon containing 35.0 L of helium at 98.0 kPa is released and rises. Assuming the temperature is constant, find out the volume of the balloon when the atmospheric pressure is 25.0 kPa.
8. A small oxygen canister contains 110 mL of oxygen gas at a pressure of 3.0 atm. This oxygen is released into a balloon with a final pressure of 2.0 atm. What is the final volume of the balloon?
9. A diving bell contains 32 kL of air at a pressure of 98 kPa at the surface. About 5 m below the surface, the pressure on the air trapped inside the bell is 150 kPa (**Figure 5**). What is the volume of air in the bell, if you assume the temperature remains the same?

Answers

6. 263 kPa
7. 137 L
8. 0.16 L
9. 21 kL

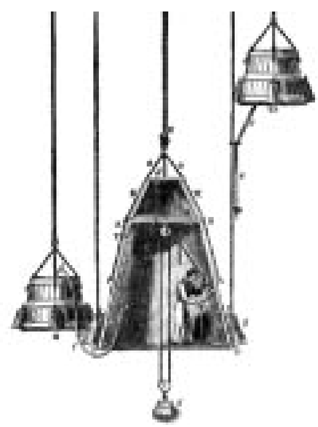

Figure 5
Before underwater diving apparatus became common, divers used a diving bell to explore underwater.

10. Why does atmospheric pressure depend on your location or vary over time at your location?

Making Connections

11. Use the Internet to investigate the invention and refinement of barometers and manometers as technologies to measure the pressure of a gas. Create a chronological flow chart or multimedia presentation including technologies, dates, names, and diagrams.

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Volume and Temperature: Charles's Law

More than a century after Boyle had determined the relationship between the volume and pressure of a gas, French physicist Jacques Charles (**Figure 6**) determined the relationship between the volume and temperature of a gas. Charles became interested in the effect of temperature on gas volume after observing the hot-air balloons that had become popular as flying machines.

Figure 6
Jacques Charles (1746–1823) designed and flew the first hydrogen balloon in 1783. His invention was based on Archimedes' concept of buoyancy, Henry Cavendish's calculations for the density of hydrogen, and his own observations. Later, his experiences and experiments led to the formulation of Charles's law.

Investigation 9.2.2

Temperature and Volume of a Gas

The purpose of this investigation is to determine how the temperature and volume of a gas are related. Complete the **Analysis** and **Evaluation** sections of the lab report.

INQUIRY SKILLS

- ○ Questioning
- ○ Hypothesizing
- ○ Predicting
- ○ Planning
- ● Conducting
- ● Recording
- ● Analyzing
- ● Evaluating
- ● Communicating

Question

What effect does increasing the temperature have on the volume of a gas?

Experimental Design

A volume of air is sealed inside a syringe, which is then placed in a water bath. As the temperature of the water (independent variable) is changed, the volume of air (dependent variable) is measured. The amount of gas inside the syringe and the pressure on the gas are two controlled variables.

(a) Read the Procedure and design a table to record your observations.

Materials

lab apron
eye protection
600-mL beaker
water, 600 mL, room temperature
plastic syringe (35–60 mL)
cap or stopper for the syringe tip
buret clamp
thermometer and clamp
ring stand
plastic stirring rod
hot plate

Heat the water slowly and ensure that the tested gas in the syringe does not eject the syringe plunger. Wear eye protection.

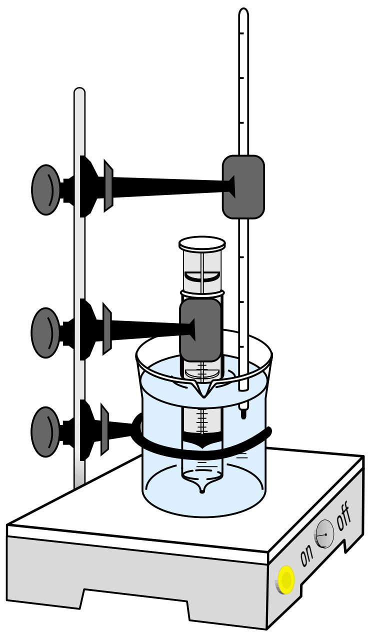

Figure 7
Setup of apparatus

Procedure

1. Set the syringe plunger to about 15–20 mL of air.
2. Seal the tip of the syringe with a cap or stopper.
3. Set up the ring stand with the 600-mL beaker on the hot plate (**Figure 7**).
4. Use the buret clamp to hold the syringe as far as possible into the beaker without touching the sides or bottom.
5. Clamp the thermometer so that the bulb is beside the end of the plunger but not touching the syringe.
6. Add room-temperature water to about 1 cm from the top of the beaker. After a few minutes, record the temperature and volume of air.
7. Turn on the hot plate.
8. Heat the water slowly, stirring occasionally.
9. Record the gas volume and temperature about every 10°C until about 90°C. (It may be necessary to tap or twist the plunger occasionally to make sure it is not stuck.)

Analysis

(b) Plot a graph of gas volume versus temperature and draw a best-fit line.
(c) According to the evidence you collected, what effect does increasing the temperature have on the volume of a gas?

Evaluation

(d) Within your lab group, discuss the Experimental Design, Materials, and Procedure used in this experiment. Decide whether these were adequate, and state your reasons in your report. List some sources of error or uncertainty and possible improvements that would raise the quality of the evidence collected.
(e) Write a summary of your discussions and indicate how certain you are about the results of this experiment.

Kelvin Temperature Scale

The mathematical equation describing the relationship between temperature and volume may not be apparent from the graph you created in Investigation 9.2.2; however, if the two variables are graphed as in **Figure 8(a)**, a straight line is obtained, so a simple relationship does exist. When the line is extrapolated downward, it meets the horizontal axis at –273°C. It appears that, if the gas did not liquefy, its volume would become zero at –273°C. If this experiment is repeated with different quantities of gas or with samples of different gases, straight-line relationships between temperature and volume are also observed. When the lines are extrapolated, they all meet at –273°C, as shown in **Figure 8(b)**. This temperature, called **absolute zero**, is the lowest possible temperature. Scientists with sophisticated technology are coming within an increasingly smaller fraction of a degree from absolute zero.

absolute zero: believed to be the lowest possible temperature

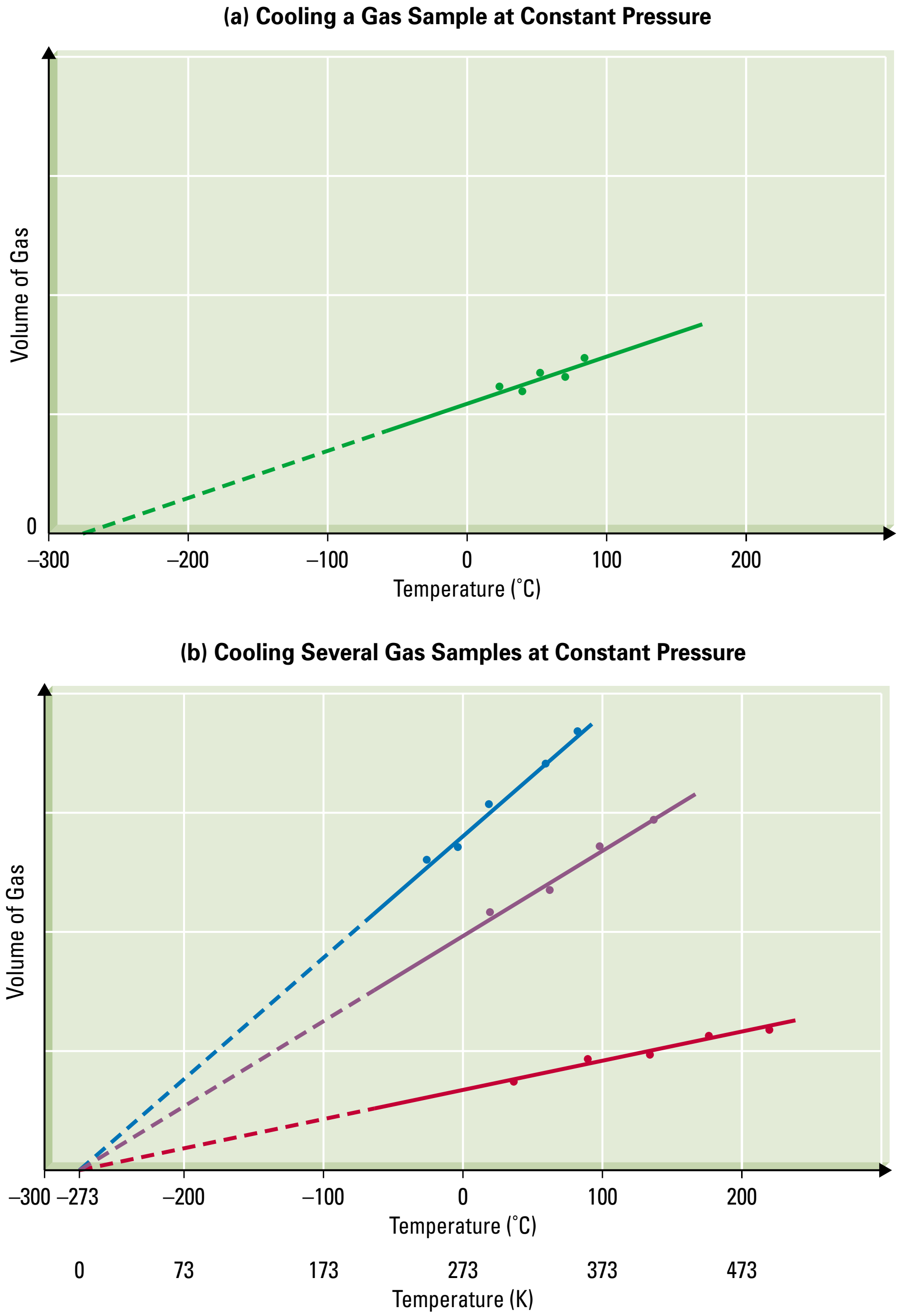

Figure 8
When the graphs of several careful volume–temperature experiments are extrapolated, all the lines meet at absolute zero, –273°C or 0 K.

DID YOU KNOW ?

Lord Kelvin

Sir William Thomson (1824–1907), also known as Lord Kelvin, was a Scottish engineer, mathematician, and physicist who profoundly influenced the scientific thought of his generation. His contributions to science included the absolute temperature scale (measured in kelvin). The style and character of Thomson's scientific and engineering work reflected his active personality. While a student at the University of Cambridge, he was awarded silver sculls for winning the university championship in racing single-seater rowing shells. He was a traveller all of his life; he spent much time in Europe and made several trips to the United States. Thomson risked his life several times during the laying of the first transatlantic cable.

Absolute zero is the basis of another temperature scale, called the absolute or **Kelvin temperature scale**. On the Kelvin scale, absolute zero (–273°C) is zero kelvin (0 K), as shown in **Figure 8(b).** (Note that no degree symbol is used for kelvin.) To convert degrees Celsius to kelvin, add 273 (**Figure 9**, page 432). STP and SATP are each defined by two exact values with infinite significant digits (i.e., STP is 273.15 K and 101.325 kPa; SATP is 298.15 K and 100 kPa). For convenience, however, use STP as 273 K and 101 kPa and SATP as 298 K and 100 kPa. Several other values are commonly rounded off to three significant digits for calculation purposes; for example, Avogadro's number, 6.02×10^{23}.

Kelvin temperature scale: a temperature scale with zero kelvin (0 K) at absolute zero and the same size divisions as the Celsius temperature scale

Answers

13. (a) 273 K
    (b) 373 K
    (c) 243 K
    (d) 298 K
14. (a) –273°C
    (b) –173°C
    (c) 27°C
    (d) 100°C

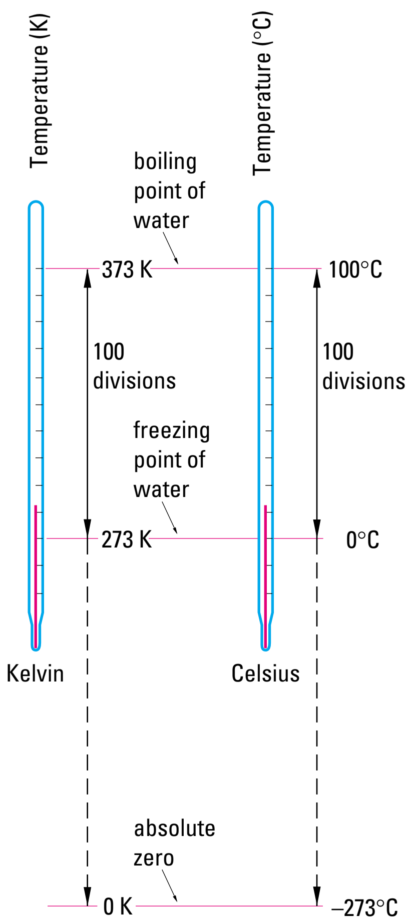

Figure 9
Jacques Charles predicted –273°C to be the temperature at which the volume of a gas would become zero, if the gas did not liquefy before reaching that temperature. Lord Kelvin considered –273°C to be the temperature at which the kinetic energy of all particles of solids, liquids, or gases would become zero. The debate continues.

Practice

Understanding Concepts

12. What is the approximate temperature for absolute zero in degrees Celsius and kelvin?
13. Convert the following Celsius temperatures to kelvin:
    (a) 0°C
    (b) 100°C
    (c) –30°C
    (d) 25°C
14. Convert the following values in kelvin to Celsius temperatures:
    (a) 0 K
    (b) 100 K
    (c) 300 K
    (d) 373 K
15. Search the Internet for research reports on how close scientists have come to reaching absolute zero. What do the reports say about whether the kinetic energy of all particles is zero at absolute zero?

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Charles's Law

The relationship in kelvin between the volume and temperature of a gas is shown in **Figure 8(b)**. This relationship is described as a direct variation; that is, as the temperature increases, the volume increases. Mathematically, this relationship is represented as

$$v = kT$$

where T represents the temperature in kelvin. This means that the quotient of the two variables (v/T) has a constant value (k), which is the slope of the straight-line graph (**Figure 8(b)**). A constant value is clearly shown by the analysis in **Table 5**.

Table 5: Analysis of Temperature and Volume of a Gas Sample

Temperature, t (°C)	Temperature, T (K)	Volume, v (L)	Constant, v/T (L/K)
25	298	5.00	0.0168
50	323	5.42	0.0168
75	348	5.84	0.0168
100	373	6.26	0.0168
125	398	6.68	0.0168

The relationship between volume and absolute temperature is known as **Charles's law**, which states that, *as the temperature of a gas increases, the volume increases proportionally, provided that the pressure and amount of gas remain constant* (**Figure 10**). Charles's law can be conveniently written comparing any two sets of volume and temperature measurements:

$$\frac{v_1}{T_1} = k \text{ and } \frac{v_2}{T_2} = k$$

$$\text{therefore, } \frac{v_1}{T_1} = \frac{v_2}{T_2} \text{ (\textit{Charles's law})}$$

This can also be expressed as a calculation of a new volume directly related to the temperatures ratio:

$$v_2 = v_1\frac{T_2}{T_1} \text{ or } v_2 = \frac{v_1 T_2}{T_1}$$

Charles's law: the volume of a gas varies directly with its temperature in kelvin, if the pressure and amount of gas are constant

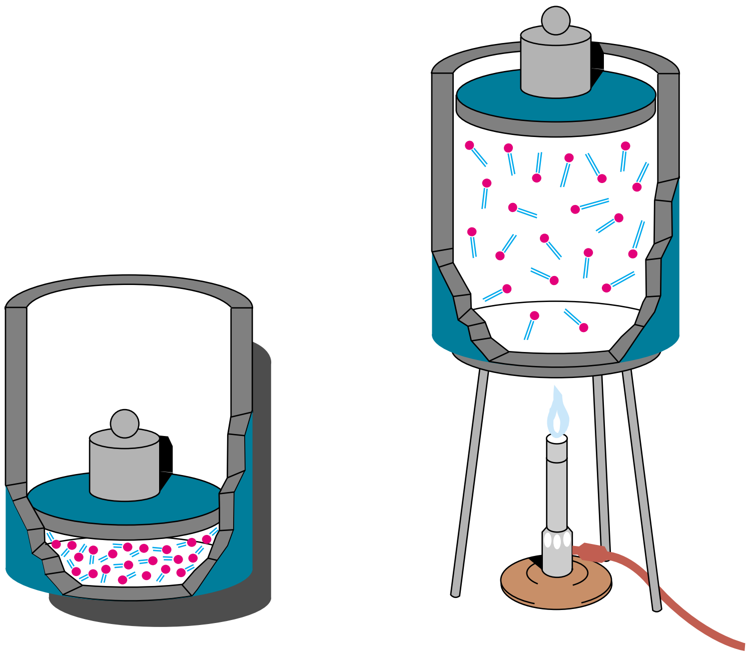

Figure 10
The volume of a gas—in this case, in a container with a movable piston—increases as the temperature of the gas increases. The pressure, equal to the pressure exerted by the mass, the piston, and the atmosphere, remains constant.

Sample Problem 3

A gas inside a cylinder with a movable piston (**Figure 10**) is to be heated to 315°C. The volume of gas in the cylinder is 0.30 L at 25°C. What is the final volume when the temperature is 315°C?

Solution

$v_1 = 0.30$ L

$T_1 = 25°C = 298$ K

$v_2 = ?$

$T_2 = 315°C = 588$ K

$$\frac{v_1}{T_1} = \frac{v_2}{T_2}$$

$$v_2 = \frac{v_1 T_2}{T_1}$$

$$= \frac{0.30 \text{ L} \times 588 \cancel{\text{K}}}{298 \cancel{\text{K}}}$$

$$v_2 = 0.59 \text{ L}$$

or $$v_{N_2} = 0.30 \text{ L} \times \frac{588 \cancel{\text{K}}}{298 \cancel{\text{K}}}$$

$$v_{N_2} = 0.59 \text{ L}$$

The final volume of the nitrogen gas at 315°C is 0.59 L.

Answers

16. 11.3 mL
17. 0.12 L
18. 26% increase
19. 1.20 L
21. (a) 3.82:1.00

Practice

Understanding Concepts

16. Butane lighters work very poorly outdoors in very cold weather. If 12.7 mL of butane gas is released from a lighter at 22°C, what volume would this same amount of butane occupy at –11°C?
17. An open, "empty" 2-L plastic pop container, which has an actual inside volume of 2.05 L, is removed from a refrigerator at 5°C and allowed to warm up to 21°C on a kitchen counter. What volume of air, measured at 21°C, will leave the container as it warms?
18. Cooking pots have loose-fitting lids to allow air to escape while food is being heated. If a 1.5-L saucepan is heated from 22°C to 100°C, any gas in the pan will increase in volume by what percentage?
19. Jacques Charles became interested in temperature–volume relationships for gases because of his curiosity about hot-air balloon flight, at a time when burning straw, with all its hazards and inconveniences, was used to heat the interior air. Hot-air balloons are open containers that maintain the air inside at (very nearly) atmospheric pressure. When a modern balloon's propane burner has warmed the air inside from an average value of 20°C to an average value of 80°C, find the final volume of each 1.00 L of air that was initially in the balloon.

Applying Inquiry Skills

20. A student decides to make a gas expansion thermometer by trapping some air (about 50–70 mL) inside an inverted 100-mL graduated cylinder, the open end of which is submerged in a beaker of water. The student reasons that she should be able to calculate the temperature of the surrounding air by measuring the volume of air inside the cylinder using the graduated scale on the cylinder walls.

    Evaluate the design of this technology, using your knowledge of gas behaviour, and predict whether this design would provide accurate values. Suggest possible improvements.

Making Connections

21. Jet aircraft engines use energy from burning fuel to power the process of taking in cold air and releasing hot gases. The 78% of the intake air that is nitrogen reacts only in negligible amounts, so basically four-fifths of the air is just heated strongly. The expanding gas mixture escapes backward, and the reaction of this force drives the engine forward.
    (a) Assuming the $N_{2(g)}$ in the air is heated from –60°C to 540°C in the engine, express the volume increase as a ratio of final volume to initial volume, to three significant digits.
    (b) Describe what other work the expanding gases in a jet engine must do, besides providing forward thrust. (Hint: An older term is "turbojet" engine.)

Pressure and Temperature Law

If you read the warning on any aerosol can, such as a can of spray paint, you will see a caution about the danger of the can exploding if heated, for example, in a fire. As you might expect from this warning, raising the temperature increases the pressure of the gas inside the can until it can no longer contain the pressure and the can ruptures. Mathematically, the direct variation is represented as

$$p = kT \quad or \quad \frac{p}{T} = k$$

If we assume the volume and amount of the gas remain the same, the quotient of the two variables (p/T) has a constant value (k). This means that two sets of pressure and temperature measurements can easily be compared. This is the **pressure and temperature law** (sometimes called Gay-Lussac's law):

$$\frac{p_1}{T_1} = \frac{p_2}{T_2} \quad \textit{(pressure and temperature law)}$$

pressure and temperature law: The pressure exerted by a gas varies directly with the absolute temperature if the volume and amount of gas remain constant.

This can also be expressed as a calculation of a new pressure directly related to the temperatures ratio:

$$p_2 = p_1 \frac{T_2}{T_1}$$

Note the similarity between this law and Charles's law comparing the volume and temperature of a gas. Both laws represent direct relationships and both require the use of absolute temperature in kelvin.

Sample Problem 4

A sealed storage tank contains argon gas at 18°C and a pressure of 875 kPa at night. What is the new pressure if the tank and its contents warm to 32°C during the day?

Solution

$T_1 = 18°C = 291\ K$
$p_1 = 875\ kPa$
$p_2 = ?$
$T_2 = 32°C = 305\ K$

$$\frac{p_1}{T_1} = \frac{p_2}{T_2}$$

$$p_2 = \frac{p_1 T_2}{T_1}$$

$$= \frac{875\ kPa \times 305\ \cancel{K}}{291\ \cancel{K}}$$

$$p_2 = 917\ kPa$$

or $p_{Ar} = 875\ kPa \times \dfrac{305\ \cancel{K}}{291\ \cancel{K}}$

$$p_{Ar} = 917\ kPa$$

The new pressure is 917 kPa.

DID YOU KNOW ?

Gay-Lussac's Law?

According to some books, the French chemist Joseph Gay-Lussac discovered the direct relationship between the temperature and pressure of a gas in the early 1800s. Therefore, this relationship is sometimes called Gay-Lussac's law. However, history of science references say that Charles, Dalton, and Gay-Lussac were all involved in investigating this relationship, with Charles and Dalton doing their work before Gay-Lussac.

Practice

Understanding Concepts

22. A closed, "empty" tank containing air at 97 kPa and 22°C survives intact in a fire. If the tank is able to withstand a maximum internal pressure of 350 kPa, what is the maximum temperature it could have reached during the fire?
23. Use the kinetic molecular theory to provide an explanation for the increase in pressure of a constant volume of gas when the temperature rises.

Answer

22. 791°C

Answers

24. (a) 167 kPa

    (b) 2.80%

25. (a) 328 kPa

    (b) 228 kPa

Applying Inquiry Skills

24. A sample of neon gas in a gas cylinder has a pressure of 125 kPa at 300 K. The cylinder is slowly heated to a temperature of 400 K.
    (a) Using the pressure and temperature law, predict the pressure at 400 K.
    (b) If the measured pressure in the cylinder at 400 K is 162 kPa, what is the percent difference?

Making Connections

25. Car manufacturers suggest that you check and adjust the air in the tires of a car when the tires are cold. To do this, you use a pressure gauge, which must read zero before you attach it to the valve stem of the tire. Therefore, the pressure reading on the gauge is actually the amount by which the tire pressure exceeds atmospheric pressure.
    (a) Suppose the pressure gauge shows a tire pressure of 210 kPa (for a total pressure of 310 kPa) at 21°C. After driving for a period of time, the tires reach a temperature of 38°C. What is the new total pressure?
    (b) What would the pressure gauge read at the tire temperature of 38°C?
    (c) What problems may be created if you set the recommended tire pressure when the tires are hot from a long period of highway driving?

The Combined Gas Law

combined gas law: The product of the pressure and volume of a gas sample is proportional to its absolute temperature in kelvin; $pv = kT$.

When Boyle's, Charles's, and the pressure–temperature laws are combined, the resulting **combined gas law** states the relationship among the volume, temperature, and pressure of any fixed amount of gas:

Boyle's law:	$pv = \text{a constant}$	(T and n are controlled variables)
Charles's law:	$\frac{v}{T} = \text{a constant}$	(p and n are controlled variables)
pressure–temperature law:	$\frac{p}{T} = \text{a constant}$	(v and n are controlled variables)

If the product pv is constant at a fixed temperature, then $p(\frac{v}{T})$ should also be a constant because v divided by a constant temperature is also constant. If the temperature changes, then Charles's law tells us that the ratio $\frac{v}{T}$ is constant at a fixed pressure. Therefore, multiplying a constant pressure by a constant ratio of volume to temperature certainly produces a number that is a constant. Alternatively, you could employ the same reasoning using Boyle's and the pressure and temperature laws to again show that $(\frac{p}{T})v$ must also be constant. Using this reasoning or a mathematical method of joint variation, we can conclude that the product of the pressure and volume of a gas divided by its absolute temperature is a constant as long as the amount of gas is controlled, that is, does not change:

$$\frac{pv}{T} = k$$

The relationship can be expressed in a convenient form for calculations involving changes in volume, temperature, or pressure for a particular gas sample:

$$\frac{p_1v_1}{T_1} = \frac{p_2v_2}{T_2} \quad (\textit{combined gas law})$$

The combined gas law is a useful starting point for all cases involving pressure, volume, and temperature, even if one of these variables is constant (as in Boyle's, Charles's, and the pressure–temperature laws). A variable that is constant can easily be eliminated from the combined gas law equation. For example, a steel cylinder with a fixed volume contains a gas at a pressure of 652 kPa and a temperature of 25°C. If the cylinder is heated to 150°C, what will be the new pressure? Because the volume is constant, we can cancel v_1 and v_2 from the combined gas law equation because $v_2 = v_1$:

$$\frac{p_1\cancel{v_1}}{T_1} = \frac{p_2\cancel{v_2}}{T_2}$$

We can now solve for p_2 and then substitute the pressures and temperatures (after converting to kelvin):

$$p_2 = \frac{p_1 T_2}{T_1}$$

$$= \frac{652 \text{ kPa} \times 423 \cancel{\text{K}}}{298 \cancel{\text{K}}}$$

$$P_2 = 925 \text{ kPa}$$

This can also be expressed as a calculation of a new pressure directly related to the temperatures ratio and inversely related to the volumes ratio (which cancels to one in this case):

$$p_{gas} = p_1 \frac{T_2}{T_1}\frac{\cancel{v_1}}{\cancel{v_2}}$$

$$= 652 \text{ kPa} \times \frac{423 \cancel{\text{K}}}{298 \cancel{\text{K}}}$$

$$p_{gas} = 925 \text{ kPa}$$

If we assume that the steel walls are sufficiently strong, the gas will have a pressure of 925 kPa inside the cylinder.

Sample Problem 5

A balloon containing hydrogen gas at 20°C and a pressure of 100 kPa has a volume of 7.50 L. Calculate the volume of the balloon after it rises 10 km into the upper atmosphere, where the temperature is –36°C and the outside air pressure is 28 kPa. Assume that no hydrogen gas escapes and that balloons are free to expand so that the gas pressure within them remains equal to the air pressure outside.

Solution

$T_1 = 20°\text{C} = 293 \text{ K}$

$p_1 = 100 \text{ kPa}$

$v_1 = 7.50 \text{ L}$

$v_2 = ?$

$T_2 = -36°\text{C} = 237 \text{ K}$

$p_2 = 28 \text{ kPa}$

$$\frac{p_1 v_1}{T_1} = \frac{p_2 v_2}{T_2}$$

$$v_2 = \frac{p_1 v_1 T_2}{p_2 T_1}$$

$$v_2 = \frac{100\ \cancel{kPa} \times 7.50\ L \times 237\ \cancel{K}}{28\ \cancel{kPa} \times 293\ \cancel{K}}$$

$$= 22\ L$$

or $$v_{balloon} = 7.50\ L \times \frac{100\cancel{kPa}}{28\cancel{kPa}} \times \frac{237\ \cancel{K}}{293\cancel{K}}$$

$$v_{balloon} = 22\ L$$

The volume of the balloon is 22 L.

Figure 11
The lightness of baked goods such as bread and cakes is a result of gas bubbles trapped in the dough or batter when it is heated. The leavening, or production of gas bubbles, can be due to vaporization of water, expansion of gases already in the dough or batter, or leavening agents such as yeast and baking powder. Yeasts are living organisms that feed on sugar, producing carbon dioxide and either water or ethanol; baking powder is a mixture of sodium hydrogen carbonate and a solid acid that react together to produce carbon dioxide; the bubbles of gas are part of the light and delectable baked goods that result from kitchen chemistry.

SUMMARY Gas Laws

STP: 0°C and 101.325 kPa (exact values)
SATP: 25°C and 100 kPa (exact values)

101.325 kPa = 1 atm = 760 mm Hg (exact values) or 101 kPa (for calculation)
absolute zero = 0 K or –273.15°C, or –273°C (for calculation)
T (K) = t (°C) + 273 (for calculation)

Boyle's law: $p_1v_1 = p_2v_2$ (for constant temperature and amount of gas)

Charles's law: $\frac{v_1}{T_1} = \frac{v_2}{T_2}$ (for constant pressure and amount of gas)

pressure–temperature law: $\frac{p_1}{T_1} = \frac{p_2}{T_2}$ (for constant volume and amount of gas)

combined gas law: $\frac{p_1v_1}{T_1} = \frac{p_2v_2}{T_2}$ (for constant amount of gas)

Practice

Understanding Concepts

26. A syringe contains 50.0 mL of a gas at a pressure of 101 kPa. The end is sealed and the plunger is pushed to compress the gas to a volume of 12.5 mL. What is the new pressure, assuming constant temperature?
27. Carbon dioxide produced by yeast in bread dough causes the dough to rise, even before it is baked (**Figure 11**). During baking, the carbon dioxide gas expands. Predict the final volume of 0.10 L of carbon dioxide in bread dough that is heated from 25°C to 190°C at a constant pressure.
28. A storage tank is designed to hold a fixed volume of butane gas at 150 kPa and 35°C. To prevent dangerous pressure buildup, the tank has a relief valve that opens at 250 kPa. At what (Celsius) temperature does the valve open?
29. A balloon has a volume of 5.00 L at 20°C and 100 kPa. What is its volume at 35°C and 90 kPa?
30. A cylinder of helium gas has a volume of 1.0 L. The gas in the cylinder exerts a pressure of 800 kPa at 30°C. What volume would this gas occupy at SATP?
31. For any of the calculations in the previous questions, does the result depend on the identity of the gas? Explain briefly.

Answers

26. 404 kPa
27. 0.16 L
28. 240°C
29. 5.8 L
30. 7.9 L

32. A 2.0-mL bubble of gas is released at the bottom of a lake where the pressure is 6.5 atm and the temperature is 10°C. What is the volume of the gas bubble when it reaches the surface, where the pressure is 0.95 atm and the temperature is 24°C?
33. What assumption was made in all of the previous calculations?

Answer

32. 14 mL

Making Connections

34. Popcorn is a favourite snack food for many people (**Figure 12**). As you learned in Chapter 4, the corn kernel is heated, and some of the moisture inside the kernel vaporizes, starting a chain of events that leads to the tasty popped corn.
    (a) If we assume a constant volume kernel (before popping), what happens to the pressure inside the kernel as the temperature increases? Justify your answer using appropriate mathematical equations or relations.
    (b) The pressure inside the kernel forces some superheated water and steam to penetrate into the starch granules, making them soft and gelatinous. When the hull of the kernel breaks at about 900 kPa, what happens to the volume of water vapour when the pressure quickly drops to about 100 kPa? Justify your answer using appropriate mathematical equations or relations.

Reflecting

35. When you are solving questions, the gas laws are like tools. How do you know which gas law is the appropriate one to use? How can the combined gas law be used instead of either Boyle's or Charles's law?

Figure 12
Popcorn was used by the original Native peoples in North America long before the arrival of Europeans. The popping method used very hot clay pots, which is a method similar to today's hot-air poppers.

Section 9.2 Questions

Understanding Concepts

1. Copy and complete **Table 6**. Show your work using appropriate conversion factors.

Table 6: Using Pressure Units

	Pressure (kPa)	Pressure (atm)	Pressure (mm Hg)
(a)		0.875	
(b)	25.0		
(c)			842

2. Copy and complete **Table 7**.

Table 7: Celsius and Kelvin

	t (°C)	*T* (K)
(a)	25	
(b)	–35	
(c)		312
(d)		208

(continued)

3. The gas laws described in this section involve the properties of volume, pressure, and temperature. Some of these variables have a direct relationship (as one increases, so does the other), and some have an inverse relationship (as one increases, the other decreases). For each pair of the following variables, state whether the relationship is direct or inverse:
   (a) pressure and volume at a constant temperature
   (b) temperature and volume at a constant pressure
   (c) temperature and pressure at a constant volume
   (d) What other property of a gas must also be constant for all of the above?
4. An automobile tire has an internal volume of 27 L at 225 kPa and 18°C.
   (a) What volume would the air inside the tire occupy if it escaped? (Atmospheric pressure at the time is 98 kPa and the temperature remains the same.)
   (b) How many times larger is the new volume compared with the original volume? How does this compare with the change in pressure?
5. In a cylinder of a diesel engine, 500 mL of air at 40.0°C and 1.00 atm is powerfully compressed just before the diesel fuel is injected. The resulting pressure is 35.0 atm. If the final volume is 23.0 mL, what is the final temperature in the cylinder?

Applying Inquiry Skills

6. The purpose of the following investigation is to test the combined gas law for the relationship between the pressure and the temperature of a gas. Complete a report, including the **Hypothesis, Experimental Design, Analysis,** and **Evaluation**.

Question

What effect does the temperature of nitrogen gas have on the pressure it exerts (**Figure 13**)?

Figure 13
This apparatus consists of a hollow metal sphere to which a pressure gauge is attached. Because the gas inside the sphere cannot expand, the relationship between temperature and pressure of a gas can be determined.

Hypothesis

(a) State the hypothesis used to answer the Question and provide your reasoning.

Experimental Design

(b) Briefly describe an experiment, using the apparatus in **Figure 13**, that would allow you to answer the Question.

Table 8: Evidence

Temperature (°C)	Pressure (kPa)
0	100
20	106
40	115
60	123
80	129
100	135

Analysis

(c) Analyze the Evidence in **Table 8**. Include in your analysis a graph and final word statement answering the Question.

Evaluation

(d) Evaluate the Evidence and the Hypothesis.

Making Connections

7. For a typical geyser (**Figure 14**), underground water seeps into a deep narrow shaft in the ground and is heated from below. Because of the depth, the pressure on the water is high so the water at the bottom of the shaft boils at a much higher temperature than normal.
   (a) What happens to the volume of a 1.0-L bubble of water vapour at 130°C and 305 kPa when it reaches the surface, where the conditions are 93°C and 101 kPa?
   (b) Why is a narrow shaft necessary to produce the geyser effect?

Figure 14
Geysers are unusual and dramatic examples of geothermal energy used to heat water in a confined space.

9.3 Compressed Gases

Not only are gases a major part of our lives, but compressed gases, that is, gases at pressures above atmospheric pressure, are particularly useful:

- The tires of vehicles contain pressurized air.
- Many people use gas barbecues with a pressurized propane fuel tank.
- Aerosol cans contain a propellant that carries the contents of the can out the nozzle; the propellant is a pressurized gas.
- Major surgery usually involves oxygen administered from a pressurized oxygen tank and is often accompanied by an anesthetic, which may also be a pressurized gas, such as dinitrogen monoxide.

Certain occupations require some work with pressurized gases. In the medical field, paramedics and doctors use oxygen tanks. Firefighters use compressed air tanks like those used by underwater divers. Some welders use oxyacetylene torches (**Figure 1**). This form of welding requires both a pressurized oxygen tank and a pressurized acetylene tank. Many scientists and their graduate students routinely use pressurized gases for research because the gas is part of the reaction system or because it provides an inert (nonreactive) environment. Noble gases, such as argon, are also used to provide an inert environment in the computer chip industry, where oxygen would cause undesirable reactions.

The chemical safety hazards of some gases are similar to those of many other chemicals, which may be corrosive, toxic, flammable, dangerously reactive, or oxidizing agents. What makes compressed gases much more dangerous is the physical hazard of a potential rocket. In commercial gas cylinders, gas pressures can be as high as 15 MPa (about 150 atm). The hole in the tank, to which the valve stem and valve are connected, is the diameter of a pencil. If the gas is suddenly released through such a small opening, the very great pressure propels the tank, making it a formidable projectile. If the tank is mishandled, dropped, or falls over and the valve stem breaks, the tank can fly through solid brick walls and cause considerable damage.

Figure 1
The use of a controlled mixture of oxygen and acetylene provides the best combustion and very high temperatures necessary for cutting or welding metal. Note the two hoses leading to the torch.

Making Connections

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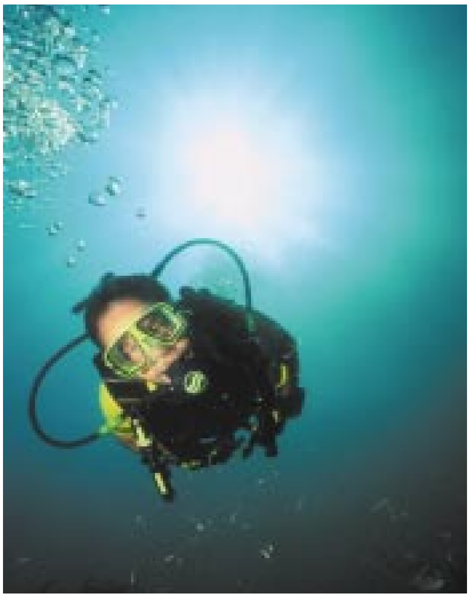

Figure 2
Every 10 m of depth adds about 1 atm (100 kPa) of pressure to the normal air pressure. At a depth of 20 m, the total pressure is about 300 kPa. In order to breathe, the air pressure then must be about 300 kPa.

Professional and recreational underwater divers face risks associated with the use of a pressurized air tank or scuba (self-contained underwater breathing apparatus). The tank containing compressed air is attached to a regulator that releases the air at the same pressure as the underwater surroundings (**Figure 2**). The pressure underwater can be quite substantial. Breathing pressurized air is necessary to balance the internal and external pressures on the chest to allow divers to inflate their lungs. However, this creates problems if divers ascend to normal pressure too quickly or while holding their breath. According to Boyle's law, if the pressure is decreased, the volume of air increases. When the volume of air is contained in the lungs, the lungs would expand to accommodate the increased volume, but this is very dangerous because the lungs could rupture. This is one reason why a person needs an understanding of gases and gas laws in order to obtain a scuba-diving licence and to dive safely.

Another problem of breathing air under pressure is that it forces more air to dissolve in the diver's bloodstream. If the diver comes up too quickly, the solubility of air (mostly nitrogen) decreases as the pressure decreases, and nitrogen bubbles form in the blood vessels. These nitrogen bubbles are the cause of a diving danger called "the bends" (so named because divers typically bend over in agony as they try to relieve the pain). Nitrogen bubbles are especially dangerous if they form in the brain or spinal cord. The bends may be avoided by ascending very slowly or corrected by using a decompression chamber.

Practice

Answer

1. (a) 18 L

Understanding Concepts

1. A scuba diver with a total lung capacity of 6.0 L breathes air at a pressure of 300 kPa and returns to the surface (100 kPa) while trying to hold her breath.
   (a) Calculate the new volume of air in her lungs at surface conditions.
   (b) Why is this a dangerous situation?
   (c) What assumption is made for the calculation in (a)?
2. Compressed gases, including those in aerosol cans, should not be heated. Use your knowledge of the properties of gases to describe the potential safety hazard.

Making Connections

3. Identify several careers that involve work with pressurized gases. How does knowledge of gas properties help those people in their jobs?
4. Suggest a chemical or physical reason why pressure regulators, which control the release of a gas from a pressurized tank, are labelled "for use only with certain gases." For example, only a regulator marked "oxygen" should be used on oxygen tanks.
5. Helium has many uses (**Figure 3**) but one that is familiar to many people is its use in party balloons. Sometimes people inhale helium because it produces an unusual change in a person's voice. In a paragraph based on your Internet research, describe this effect as well as the dangers of inhaling helium.

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Figure 3
Modern airships are filled with the light noble gas, helium. Helium is also used in party balloons.

Section 9.3 Questions

Making Connections

1. Identify two consumer or commercial products that use or contain compressed air and two that involve other compressed gases.
2. How is the safety hazard of compressed gases different from that of other chemicals? State the potential hazard, how it may be created, and some possible consequences.
3. Even if a compressed gas is slowly released at a normal pressure, the gas may still be dangerous. What are some of the possible dangers?

DID YOU KNOW ?

First to Commercially Produce Helium

John McLennan (1867–1935), University of Toronto research scientist and inventor, was the first to produce commercial quantities of helium (by extracting it from Alberta natural gas). The process dropped the price from \$20 000/m^3 to \$3/m^3. Helium is a safe alternative to hydrogen for airships.

9.4 The Ideal Gas Law

In this book, all real gases are dealt with in calculations as if they were ideal. An **ideal gas** is a hypothetical gas that obeys all the gas laws perfectly under all conditions; that is, it does not condense into a liquid when cooled, and graphs of its volume and temperature and of its pressure and temperature relationships are perfectly straight lines. A single, ideal gas equation describes the relationship among pressure, temperature, volume, and amount of matter, the four variables that define an ideal gaseous system.

ideal gas: a hypothetical gas composed of particles that have zero size, travel in straight lines, and have no attraction to each other (zero intermolecular force)

- According to Boyle's law, the volume of a gas is inversely proportional to the pressure: $v \propto 1/p$.
- According to Charles's law, the volume of a gas is directly proportional to the absolute temperature: $v \propto T$.
- According to the pressure–temperature law, the pressure of a gas is directly proportional to the absolute kelvin temperature: $p \propto T$.
- As anyone knows who has blown up a balloon, the more air you blow into the balloon, the bigger it gets; in other words, the greater the amount of air, in moles, at the same temperature and pressure, the greater the volume. Therefore, $v \propto n$.

Combining the three mathematical relationships for the volume of a gas produces the following relationship:

$$v \propto \frac{1}{p} \times T \times n$$

Another way of stating this is

$$v = (\text{constant, } R) \times \frac{1}{p} \times T \times n$$

$$v = \frac{nRT}{p}$$

$$pv = nRT$$

This last equation is known as the **ideal gas law**, and the constant R is known as the **gas constant**. The value for the gas constant is obtained by substituting known values into the ideal gas law and solving for R. For example, references state that at STP, 1.00 mol of an ideal gas would occupy a volume of 22.414 L.

ideal gas law: The product of the pressure and volume of a gas is directly proportional to the amount and the kelvin temperature of the gas; $pv = nRT$.

gas constant: the constant of variation, R, that relates the pressure in kilopascals, volume in litres, amount in moles, and temperature in kelvin of an ideal gas

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gas constant: the constant of variation, R, that relates the pressure in kilopascals, volume in litres, amount in moles, and temperature in kelvin of an ideal gas

DID YOU KNOW ?

van der Waals Forces

In 1873, Johannes Diderik van der Waals (1837–1923) hypothesized the existence of attraction between gas molecules to explain deviations from the ideal gas law. The general attraction, called van der Waals forces, was later explained as being London forces by Fritz London and dipole–dipole forces by, for example, Peter Debye.

$v = 22.414$ L

$n = 1.00$ mol (to give a certainty for R of three significant digits)

$p = 101.325$ kPa

$T = 0°C = 273.15$ K

$$pv = nRT$$

$$R = \frac{pv}{nT}$$

$$= \frac{101.325 \text{ kPa} \times 22.414 \text{ L}}{1.00 \text{ mol} \times 273.15 \text{ K}}$$

$$= \frac{8.31 \text{ kPa}\cdot\text{L}}{\text{mol}\cdot\text{K}}$$

As with several other constant values used for calculations in this book, a certainty of three significant digits for R is normally used. The value of the gas constant depends on the units chosen to measure volume, pressure, and temperature. If any three of the four variables in the ideal gas law are known, the fourth can be calculated by means of this equation. However, often the mass of a gas is a known quantity. In this case, a two-step calculation is required.

For example, 0.78 g of hydrogen at 22°C and 125 kPa is produced. What volume of hydrogen would be expected? To use the ideal gas law, you first need to convert the mass into an amount in moles of hydrogen:

$$n_{H_2} = 0.78 \cancel{\text{g}} \times \frac{1 \text{ mol}}{2.02 \cancel{\text{g}}} = 0.39 \text{ mol}$$

(Retain the unrounded value for subsequent calculation.) Now you can use the ideal gas law to determine the volume of hydrogen at the conditions specified:

$$pv = nRT$$

$$v_{H_2} = \frac{nRT}{p}$$

$$= \frac{0.39 \cancel{\text{mol}} \times \frac{8.31 \cancel{\text{kPa}} \cdot \text{L}}{1 \cancel{\text{mol}} \cdot \cancel{\text{K}}} \times 295 \cancel{\text{K}}}{125 \cancel{\text{kPa}}}$$

$$v_{H_2} = 7.6 \text{ L}$$

Sample Problem 1

What mass of neon gas should be introduced into an evacuated 0.88-L tube to produce a pressure of 90 kPa at 30°C?

Solution

$v = 0.88$ L

$p = 90$ kPa

$m_{Ne} = ?$

$T = 30°C = 303$ K

$$pv = nRT$$

$$n_{Ne} = \frac{pv}{RT}$$

$$= \frac{90\ \cancel{kPa} \times 0.88\ \cancel{L}}{\frac{8.31\ \cancel{kPa} \cdot \cancel{L}}{mol \cdot \cancel{K}} \times 303\ \cancel{K}}$$

$$= 0.031\ mol$$

$$m_{Ne} = 0.031\ \cancel{mol} \times \frac{20.18\ g}{1\ \cancel{mol}}$$

$$m_{Ne} = 0.63\ g$$

The mass of neon gas required to produce a pressure of 90 kPa at 30°C is 0.63 g.

Practice

Understanding Concepts

1. List three ways of reducing the volume of gas in a shock absorber (cylinder and piston) of an automobile.
2. Under what conditions is a gas closest to the properties of an ideal gas? Why?
3. What amount of methane gas is present in a sample that has a volume of 500 mL at 35.0°C and 210 kPa?
4. What volume does 50 kg of oxygen gas occupy at a pressure of 150 kPa and a temperature of 125°C?
5. The density of a gas is the mass per unit volume of the gas in units of, for example, grams per litre. By finding the mass of one litre (assume 1.00 L) of gas, you can then calculate the density of the gas.
   (a) What is the density of propane, $C_3H_{8(g)}$, at 22°C and 96.7 kPa?
   (b) If the density of air at this temperature is 1.2 g/L, what happens to propane gas that may leak from a propane cylinder in a basement or from the tank of an automobile in an underground parkade? Why is this a problem?
6. Determining the molar mass of gases is an important experiment for qualitative analysis. Starting with the ideal gas law, derive a formula to calculate the molar mass, M, of a gas, given the mass and volume of the gas at a specific pressure and temperature, and that $n = \frac{m}{M}$.

Applying Inquiry Skills

7. The purpose of the following investigation is to use the ideal gas law to determine an important property of a substance, its molar mass. Complete the **Analysis** section of the following investigation report.

   Question

   What is the molar mass of an unknown gas?

   Experimental Design

   A measured mass of the gas is collected and the volume, temperature, and pressure of the gas are measured.

Answers

3. 41.0 mmol
4. 34 kL (or 34 m^3)
5. (a) 1.74 g/L

Answers

7. (a) 28.4 g/mol
9. (a) 1.78 g/L
   (c) 1.1 g/L

Evidence

mass = 1.25 g   pressure = 100 kPa

volume = 1.00 L   temperature = 0°C

Analysis

(a) What is the molar mass of the gas?

Figure 1
A particular hot-air balloon has a fixed size when fully inflated and is open at the bottom. The air is usually heated with a propane burner.

Making Connections

8. Hot-air balloons (**Figure 1**) rise up through the air because the density of the air inside the balloon is less than the density of the outside air. Using the ideal gas law, describe how this occurs.
9. Knowledge of the densities of gases compared to air (at 1.2 g/L) can save your life.
   (a) What is the density of carbon dioxide gas at SATP?
   (b) If you are caught in a fire, is the suffocating carbon dioxide gas found closer to the floor or to the ceiling?
   (c) What is the density of carbon monoxide gas at 20°C and 98 kPa in a home?
   (d) Where should a carbon monoxide detector be located, close to the floor or close to the ceiling?
   (e) If potentially lethal carbon dioxide comes from a fire and the carbon monoxide comes from the furnace, what other variable might affect the density of these gases released within a home?

Reflecting

10. In conversations and discussions, it is not unusual for people to say, "In an ideal world, ..." followed by some statement like, "the buses would always run on time." How does the use of the word "ideal" compare with the use of this word in the "ideal gas law"? What other concepts have you learned in science that are also ideal?

INQUIRY SKILLS

- ○ Questioning
- ○ Hypothesizing
- ● Predicting
- ○ Planning
- ● Conducting
- ● Recording
- ● Analyzing
- ● Evaluating
- ● Communicating

Investigation 9.4.1

Determining the Molar Mass of a Gas

The molar mass of a compound is an important constant that, in some cases, can help to identify a substance. Molar masses of known compounds are obtained using accepted atomic molar masses from the periodic table. The purpose of this investigation is to use the ideal gas law to evaluate the Experimental Design by determining the molar mass of a well-known substance. Butane, $C_4H_{10(g)}$, is suggested, but you may substitute another gas. Complete the **Prediction, Analysis, Evaluation** and **Synthesis** sections of the lab report.

Butane is highly flammable. Do not conduct this experiment near an open flame. Flints must be removed from butane lighters.

Good ventilation in the laboratory is essential.

Eye protection and a lab apron are required.

Question

What is the molar mass of butane?

Prediction

(a) According to the chemical formula and atomic molar masses in the periodic table, what is the molar mass of butane?

Experimental Design

A sample of butane gas from a lighter is collected in a graduated cylinder by downward displacement of water. The volume, temperature, and pressure of the gas are measured, along with the change in mass of the butane lighter. The Experimental Design is evaluated on the basis of the accuracy of the experimental value for the molar mass of butane, which is compared with the accepted value.

Materials

lab apron
eye protection
butane lighter (with flint removed) or butane cylinder with tubing
plastic bucket
100–500-mL graduated cylinder
balance
thermometer
barometer

Procedure

1. Determine the initial mass of the butane lighter.
2. Pour water into the bucket until it is two-thirds full. Then completely fill the graduated cylinder with water; invert the graduated cylinder in the bucket. Ensure that no air has been trapped in the cylinder.
3. Hold the butane lighter in the water under the graduated cylinder (**Figure 2**) and release the gas until you have collected half to three-quarters of the cylinder. Make sure all the bubbles enter the cylinder.
4. Equalize the pressures inside and outside the cylinder by adjusting the position of the cylinder until the water levels inside and outside the cylinder are the same.
5. Read the measurement on the cylinder and record the volume of gas collected. Record the ambient (room) temperature and pressure. Record your measurements as precisely as possible with the equipment available.
6. Thoroughly dry the butane lighter and determine its final mass.
7. Release the butane gas from the cylinder in a fume hood or outdoors.

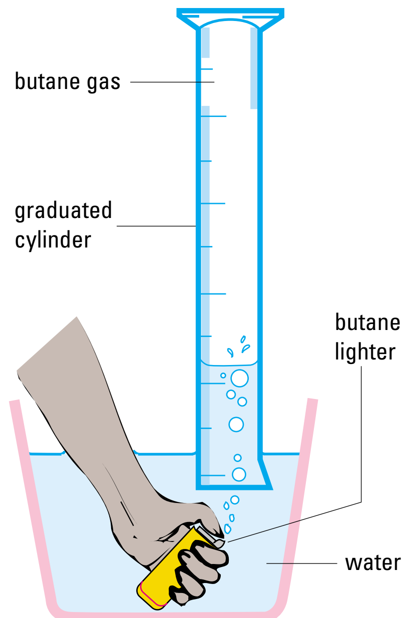

Figure 2
The gas from a butane lighter can be collected by downward displacement of water. This apparatus can be used to determine the molar mass of butane.

Analysis

(b) Using the difference in the masses of the butane container before and after the experiment, calculate the mass of butane that was released from the container.
(c) Combine the $pv = nRT$ and $n = \frac{m}{M}$ to solve for molar mass, M.
(d) Using your answers to (b) and (c), answer the Question by calculating the molar mass of butane according to your Evidence.

Evaluation

(e) Calculate the percentage difference. How accurate was your experimental answer for the molar mass of butane?
(f) Every experiment has some experimental errors or uncertainties. In order to judge whether your percentage difference was reasonable for this particular experiment, you need to consider possible flaws, experimental errors or uncertainties, and any other limitations that may affect the accuracy of your result. Within your lab group, make a list and rank the items from most to least significant.

(g) In your opinion, can the percentage difference you calculated in (e) be reasonably explained by your list of experimental errors or uncertainties? If not, explain briefly.
(h) Are the Experimental Design, Materials, and Procedure adequate for this experiment? Provide reasons.

Synthesis

(i) Evaluating the evidence from an experiment is not an easy task, and it is possible to omit a source of experimental error. In this experiment, you may or may not have realized that the gas you collected is not completely pure butane. It also contains a small quantity of water vapour. For a typical room temperature, this would correspond to about 2.6 kPa of the pressure you recorded. Subtract this value from the pressure you used, and use the new pressure to recalculate the molar mass and the percentage difference. Is your answer significantly more accurate?
(j) Are there any other possible sources of error? For example, what assumption must be made about the butane from the lighter or canister you used?

SUMMARY Properties of an Ideal Gas

- v-T and p-T graphs are perfectly straight lines
- gas does not condense to a liquid when cooled
- gas volume = 0 at absolute zero
- $pv = nRT$
- gas particles are point size (volume of particle = 0)
- gas particles do not attract each other

Section 9.4 Questions

Understanding Concepts

1. Unlike an ideal gas, a real gas condenses to a liquid when the temperature is low enough. What does this indicate about the interaction between the particles and why the gas is real versus ideal?
2. What are three variables that can determine the pressure of a gas? How is each of these variables related to the pressure?
3. Determine the pressure in a 50-L compressed air cylinder if 30 mol of air is present in the container, which is heated to 40°C.
4. At what temperature does 10.5 g of ammonia gas exert a pressure of 85.0 kPa in a 30.0-L container?
5. Using atmospheres as the pressure unit, determine the value of the gas constant, *R,* at SATP for 1.00 mol of a gas that occupies 24.8 L.
6. A 1.49-g sample of a pure gas occupies a volume of 981 mL at 42.0°C and 117 kPa.
   (a) Determine the molar mass of the compound.
   (b) If the chemical formula is known to be XH_3, identify the element "X."

9.5 Air Quality

Air quality, like water quality, has often been taken for granted. Increasingly, both have come under public scrutiny as the quality of these two requirements of life has deteriorated. The Ontario Medical Association estimates that 1900 people die prematurely each year in Ontario due to poor air quality, and almost $10 billion a year is spent in health-care costs, lost work time, and other quantifiable expenses resulting from poor air quality. In southern Ontario, the number of "smog alert days" has increased dramatically over the past few years, and the recorded asthma rate in children has increased by 60% in the past decade. Young children, the elderly, and people with respiratory illnesses are particularly at risk. Public opinion polls commissioned by the Ontario Clean Air Alliance in March 1999 in each of the Hamilton, London, and Kitchener–Waterloo regions showed that approximately 70% (of the 400 adult respondents) were concerned about air pollution, and 80% thought that the problem would continue to get worse in the future. Less than 10% gave the Ontario government a positive rating for its efforts in controlling air pollution, and more than 85% were willing to pay higher utility bills to lower emissions from coal-fired electric generating plants. Although industry, particularly the electric power plants, is a major source of air pollution, the transportation choices made by citizens are also a significant part of the air pollution problem.

Although the composition of air is well known (**Table 1**), its chemistry is very complex and not completely understood. Many of the chemical processes taking place in the atmosphere are induced by solar radiation and are intertwined with chemical processes—both naturally occurring and as a result of human activities. However, it is becoming apparent that the human activities are producing dangerous levels of air pollutants, particularly in large urban areas. Primary air pollutants include

- gases such as various nitrogen oxides, sulfur dioxide, and carbon dioxide;
- vapours from volatile organic compounds (VOCs);
- heavy metals and other toxic and carcinogenic substances emitted primarily during the combustion of coal.

An important secondary air pollutant is ground-level ozone formed when nitrogen oxides and VOCs react in sunlight. According to the Canadian Council of Ministers of the Environment (CCME), the Windsor–Quebec City corridor has one of the worst ozone problems in the country. This section focuses on some of the chemistry of nitrogen and VOCs in the atmosphere and their role in the formation of ground-level ozone.

Table 1: Components of Dry Air at Sea Level

	Composition	
Gas*	**Volume (%)**	**Volume (ppm)**
$N_{2(g)}$	78.08	7.808×10^5
$O_{2(g)}$	20.95	2.095×10^5
$Ar_{(g)}$	0.934	9.34×10^3
$CO_{2(g)}$	0.036	3.6×10^2
$Ne_{(g)}$	0.001 818	18.18
$He_{(g)}$	0.000 524	5.24
$CH_{4(g)}$	0.000 2	2
$Kr_{(g)}$	0.000 114	1.14

* Water is excluded from the table because its concentration can vary widely from one place to another.

The Nitrogen Cycle in Nature

Nitrogen is an important element in living systems (**Figure 1**, page 450). Before atmospheric nitrogen can be absorbed by plants, it must be converted into compounds suitable for assimilation. The conversion of atmospheric nitrogen into nitrates by various microorganisms in the soil (nitrogen-fixing bacteria) and in the roots of certain plants (legumes, e.g., beans and peas) is called nitrogen fixation. Lightning is another natural source of nitrogen compounds for plants. At the high temperature of a lightning strike, nitrogen and oxygen react to produce nitrogen oxides, which react with water to form nitrous and nitric acids. These acids are converted to nitrites and nitrates in the soil. Decaying plants and animals and animal waste products are in turn acted on by bacteria, and the nitrogen in them is again made available for circulation as nitrogen in the atmosphere.

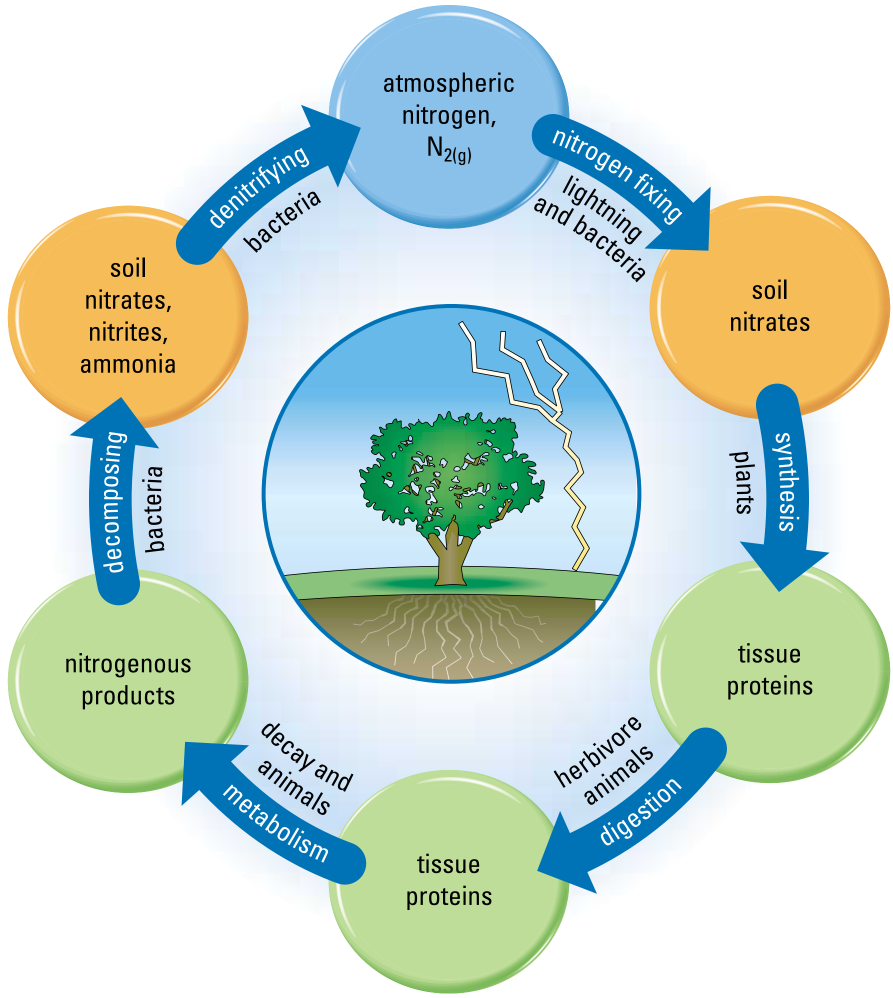

Figure 1
The nitrogen cycle is an essential link in the biosphere, the part of Earth that supports life. The proportion of nitrogen in the atmosphere remains relatively constant due to its circulation by living organisms.

Ground-Level Ozone

Ozone, $O_{3(g)}$, near the surface (as opposed to the ozone layer high in the atmosphere) is a component of smog and one of the most important health concerns of polluted air. Ozone is produced near the surface when nitrogen oxides and VOCs react in the presence of sunlight. The human activities of transportation, generating electricity, and industrial combustion have greatly increased the amount of nitrogen oxides, NO_x (e.g., $N_2O_{(g)}$, $NO_{(g)}$, $NO_{2(g)}$, $N_2O_{4(g)}$), produced annually in Canada. In 1995, transportation accounted for about 58% of NO_x, energy production 12%, and industries 24%. Total emissions of NO_x in Ontario were 538 t in 1995, rising to 565 t in 1998 (CCME Annual Progress Report, 1999). In August 2000, Ontario's Anti-Smog Action Plan committed the province to a 45% reduction of NO_x and VOCs by the year 2010. Although this sounds promising, many environmental groups seriously question whether actual reductions will occur. For example, Ontario Power Generation (formerly Ontario Hydro) has already claimed a reduction of 19 kt of NO_x when in fact its emissions have increased. The claimed reduction is possible when Ontario Power Generation buys emission credits from another industry that has earned these credits. Ways to earn credits are by improving energy efficiency and by adding pollution-reduction technologies. It is possible to gain credits by expanding operations without increasing NO_x emissions (i.e., the NO_x emissions don't actually have to be reduced).

The production of nitrogen oxides from human activities begins with the combustion of nitrogen. Nitrogen monoxide, $NO_{(g)}$, is produced whenever fuel is burned in air at a high temperature, for example, in an internal combustion engine. The higher the flame temperature, the more $NO_{(g)}$ is produced:

$$N_{2(g)} + O_{2(g)} \rightarrow 2\ NO_{(g)}$$

When released into the atmosphere, the colourless nitrogen monoxide rapidly reacts with oxygen to form nitrogen dioxide, a reddish-brown gas with an irritating odour. The brownish colour of the air in a smog-bound city is due to $NO_{2(g)}$ (**Figure 2**):

$$2\ NO_{(g)} + O_{2(g)} \rightarrow 2\ NO_{2(g)}$$

colourless reddish-brown

Figure 2
The formation of ground-level ozone requires substantial automobile traffic emitting nitrogen monoxide and VOCs, warmth, ample sunlight, and relatively little movement of the air mass.

In addition to being a toxic chemical, $NO_{2(g)}$ plays a central role in the production of ground-level ozone. Sunlight at short wavelengths (ultraviolet region) causes the decomposition of $NO_{2(g)}$:

$$NO_{2(g)} \xrightarrow{\text{UV light}} NO_{(g)} + O_{(g)}$$

The atomic oxygen formed is highly reactive and reacts with molecular oxygen to form ozone:

$$O_{(g)} + O_{2(g)} \rightarrow O_{3(g)}$$

Volatile organic compounds also contribute to the formation of ground-level ozone. Hydrocarbons containing carbon–carbon double bonds (C$=$C) are the most reactive VOCs. Once again, human activity is a significant source of this pollutant, with transportation, industrial processes, and commercial activities the main contributors. Although trees are by far the largest source of VOCs in Canada, human activities produce high concentrations of VOCs in relatively small, localized environments. This is where the problems are created. For example, when gasoline is pumped into a car, the smell of gasoline is evident and "wavy lines" are often visible in the air at the pump nozzle; this indicates that gasoline vapours are in the air. Other sources of VOCs include any products or paints that use solvents that evaporate, wood burning, dry cleaners, paint shops, and various industries (**Figure 3**, page 452).

Ozone is produced by a series of complex reactions involving VOCs, NO, and oxygen under the influence of sunlight. In summary,

$$\text{VOC} + NO_{(g)} + O_{2(g)} \rightarrow O_{3(g)} + HNO_{3(aq)} + \text{oxidized organics}$$

Although formed in urban areas, the pollutants NO_x, VOCs, and $O_{3(g)}$ can travel considerable distances in the atmosphere. This causes problems because they pollute regions that do not have high emissions or further aggravate areas that are already suffering from excess emissions, such as the Windsor–Quebec City corridor.

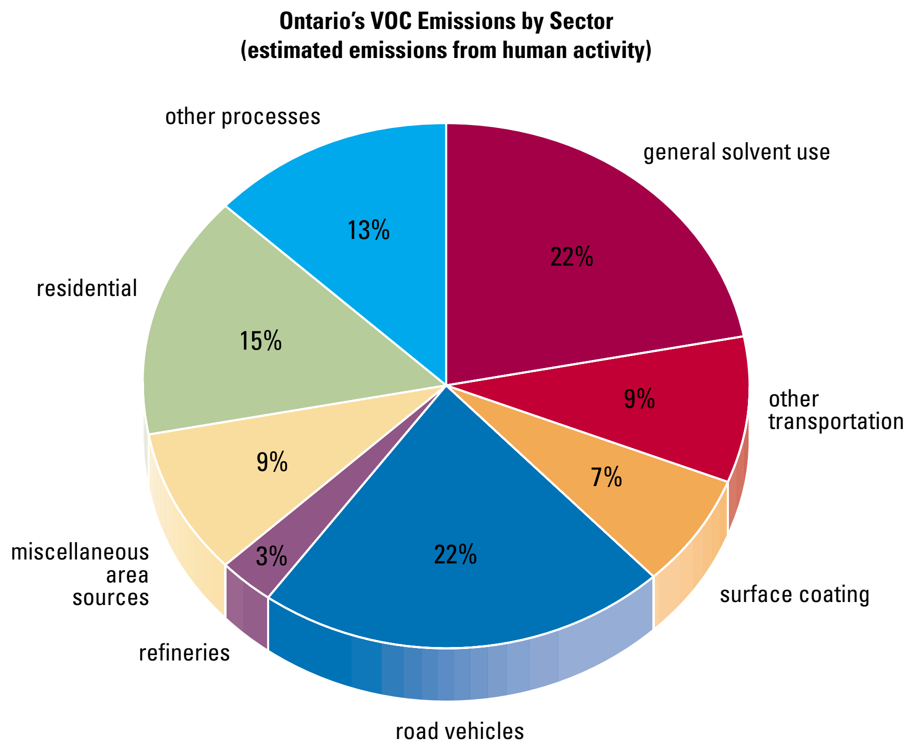

Figure 3
Transportation modes account for approximately 31% of the total VOC emissions in Ontario.

Source: Ontario Ministry of the Environment, 2000

Ozone irritates the respiratory tract and eyes, and people with heart or lung problems are at risk of lung damage. Ground-level ozone is also undesirable because it damages plants, bleaches colour from fabrics, and hardens rubber, thereby shortening the life span of tires.

Plants are damaged in several ways. According to current research, ozone causes injury to foliage, increases susceptibility to diseases in plants and trees, reduces yields in sensitive crops, and increases the mortality rates of individual trees. As a result of ozone pollution from the American Midwest, certain crops, such as white beans, can no longer be grown successfully in southwestern Ontario. The maximum allowable ozone concentration in Canada is 82 ppb averaged over a one-hour period, but this level is often exceeded in urban centres (**Figure 4**). The federal and provincial governments have recently agreed to a target of 65 ppb, averaged over an eight-hour period, by 2010.

In order to reduce the formation of ground-level ozone, it is necessary to reduce the concentration of the reactants in the air. Since it is usually the concentration of NO_x that controls the overall rate of the ozone-producing reactions, most research has focused on developing technologies to reduce the amounts of NO_x released into the atmosphere. Catalytic converters significantly reduce the level of NO_x emitted by automobile engines. In a catalytic converter the hot exhaust gases pass over a surface containing beads of rhodium, platinum, and palladium, which reduces $NO_{(g)}$ to $N_{2(g)}$ and oxidizes $CO_{(g)}$ to $CO_{2(g)}$. The air in Los Angeles used to reach ozone levels of 680 ppb but has been reduced to around 300 ppb, thanks in large part to the development of catalytic converters.

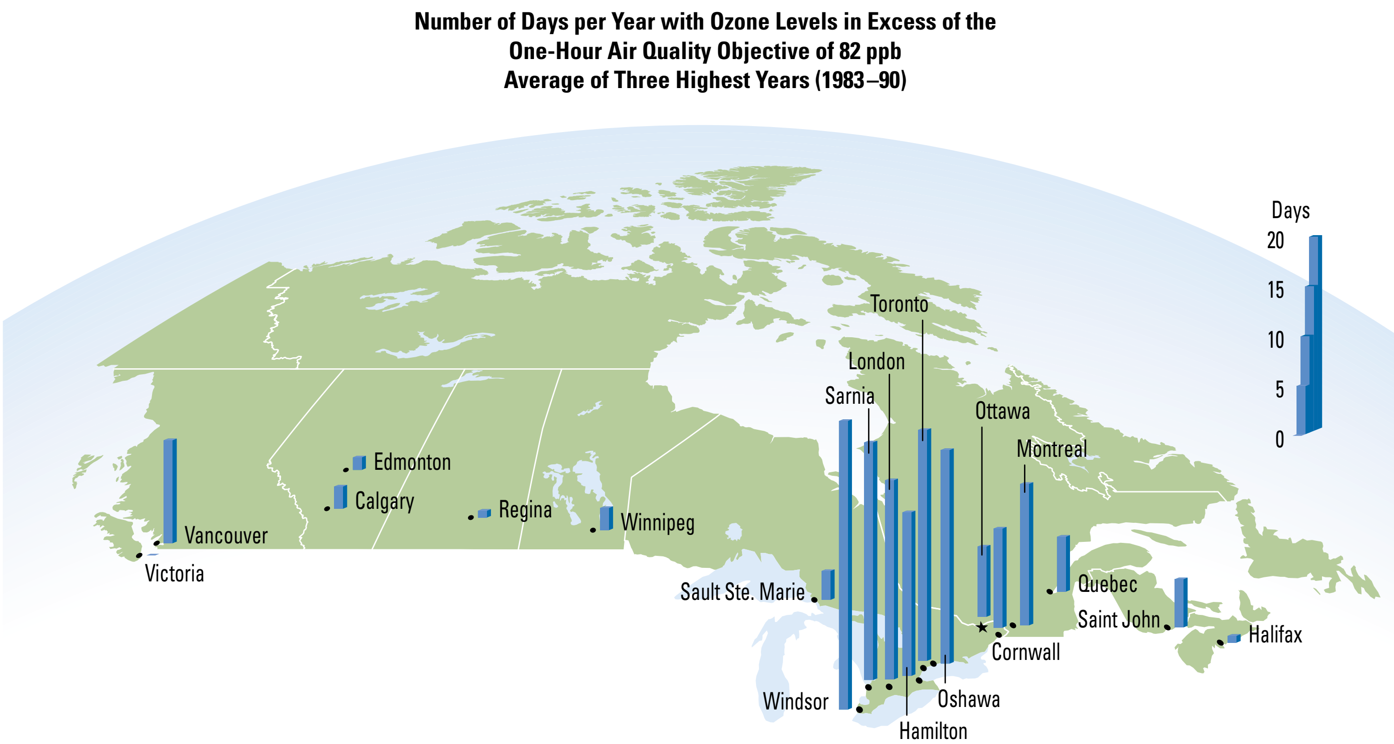

Figure 4
This 1994 Environment Canada map shows the number of days per year with ozone levels in excess of the 82 ppb current standard.

Source: Environment Canada, 1994

Practice

Understanding Concepts

1. Why is water vapour not listed in **Table 1**?
2. Describe the nitrogen cycle.
3. Which human activities contribute the most nitrogen oxides to the atmosphere?
4. Describe the role of nitrogen oxides in producing ground-level ozone.
5. An empty classroom contains 240 kL of air at SATP. Assume the air is dry.
   (a) Calculate the amount (in moles) of nitrogen, oxygen, argon, and carbon dioxide in the room.
   (b) What is the mass of each of the gases in the room? If the total mass of air was liquefied, could you carry this mass of air?
   (c) In what way will the amount of each gas change after a class has been in the room for an hour? State any assumption(s) that you make when answering this question.
   (d) At SAPT, respiration in a classroom converts about 400 L of oxygen per student per school day to carbon dioxide. If there are 30 students in the classroom for this period of time and the air is stagnant, what percentage of the oxygen is consumed?
6. How is it possible for an industry to claim a reduction of a pollutant when its actual emissions have increased?

Answers

5. (a) 7.57 kmol $N_{2(g)}$, 2.03 kmol $O_{2(g)}$, 90.5 mol $Ar_{(g)}$, and 3.5 mol $CO_{2(g)}$
   (b) 212 kg $N_{2(g)}$, 65.0 kg $O_{2(g)}$, 3.62 kg $Ar_{(g)}$, and 0.15 kg $CO_{2(g)}$
   (d) 23.9%

Making Connections

7. Weather stations and government agencies often report air quality using an air quality index (AQI). Use the Internet to discover and report on what this index includes. What is the scale that is reported in the media, including possible health warnings and effects?

   Follow the links for Nelson Chemistry 11, 9.5.

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Explore an Issue

Take a Stand: How Can We Improve the Air Quality in Our Communities?

DECISION-MAKING SKILLS

- Define the Issue
- Identify Alternatives
- Research
- Analyze the Issue
- Defend a Decision
- Evaluate

Working in small groups, research the air quality in your community. Your research should include the following:

- the major air pollutants;
- the sources of these pollutants;
- the measures currently in place to reduce the pollutants.

Follow the links for Nelson Chemistry 11, 9.5.

GO TO www.science.nelson.com

(a) Draft a letter to the appropriate department in your municipality, recommending improvements where you think they are needed, or

(b) Draft a pamphlet for the public that outlines the concerns of your group.

Section 9.5 Questions

Understanding Concepts

1. Identify the four most abundant gases in the atmosphere.
2. Identify the three most abundant noble gases in the atmosphere.
3. Water in the air is so variable that it is often not listed as a component or considered as affecting the properties of air. For example, is moist air denser than dry air? Provide your reasoning. (No calculations are necessary.)
4. List and describe the three main categories of primary air pollutants.
5. Why is ozone not classified as a primary air pollutant?
6. Describe how catalytic converters in automobiles improve air quality.
7. Air quality is the responsibility of everyone on our planet. What are some consumer, commercial, and/or industrial strategies to reduce $NO_{x(g)}$ and VOC emissions?

Chapter 9 Summary

Key Expectations

Throughout this chapter, you have had the opportunity to do the following:

- Explain different states of matter in terms of the forces among atoms, molecules, and ions. (9.1)
- Describe the gaseous state, using the kinetic molecular theory, in the terms of degree of disorder and the types of motion of atoms and molecules. (9.1)
- Describe natural phenomena and technological products associated with gases. (9.1, 9.2, 9.3, 9.4)
- Determine through experimentation the quantitative and graphical relationships among the pressure, volume, and temperature of an ideal gas. (9.2)
- Solve quantitative problems involving the following gas laws: Charles's law, Boyle's law, the combined gas law, the pressure and temperature law (Gay-Lussac's law), and the ideal gas law. (9.2, 9.4)
- Describe the quantitative relationships that exist among the following variables for an ideal gas: pressure, volume, temperature, and amount of substance. (9.2, 9.4)
- Use and interconvert appropriate units to express pressure and temperature. (9.2, 9.3, 9.4)
- Use appropriate scientific vocabulary to communicate ideas related to gases. (9.2, 9.3, 9.4)
- Identify technological products and safety concerns associated with compressed gases. (9.3)
- Identify the major and minor components of the atmosphere and describe Canadian initiatives to improve air quality. (9.5)

Key Terms

absolute zero
atmospheric pressure
Boyle's law
Charles's law
combined gas law
gas constant
ideal gas
ideal gas law
Kelvin temperature scale
kinetic molecular theory
pressure
pressure and temperature law

Make a Summary

1. Draw a simple model of a solid, a liquid, and a gas. With each model, list the
   (a) empirical properties
   (b) possible forces present
   (c) type of motion of the particles
   (d) degree of order
2. Sketch a series of graphs showing the relationship between volume and each of the following variables: pressure, temperature, and amount of gas. For each graph, include a mathematical equation and indicate the variables that are controlled.
3. Write the ideal gas law. Using arrows, label each symbol in the equation with the name of the variable and the typical units.

Reflect on your Learning

Revisit your answers to the Reflect on Your Learning questions, at the beginning of this chapter.

- How has your thinking changed?
- What new questions do you have?

Chapter 9 Review

Understanding Concepts

1. Chlorine, bromine, and iodine are all members of the halogen family and have similar chemical properties. What does the state (solid, liquid, or gas) of these elements at SATP reveal about the strength of their intermolecular forces?
2. All three states of matter for H_2O are always present on Earth: Most of the surface of Earth is covered by water and ice (**Figure 1**), and the atmosphere contains water vapour.
   (a) Which physical properties of water and ice are similar?
   (b) Which physical properties of water and ice are different?
   (c) Why is a gas like water vapour highly compressible?
   (d) In general terms, describe the degree of disorder of the H_2O molecules in all three states.
   (e) Briefly describe all three states of matter, using forces and motion of H_2O molecules.
   (f) Water vapour is a small and variable part of the composition of the atmosphere. List the two major components and two minor components of the atmosphere.
   (g) Water is a small component of the atmosphere, but it is a very important one. List some effects that water has as a component of Earth's atmosphere.

Figure 1

3. State each of the following laws in a sentence beginning, "The volume of a gas sample ..."
   (a) Boyle's law
   (b) Charles's law
   (c) the ideal gas law
4. Predict how the volume of a given mass of gas will differ when the following changes in the temperature and pressure are made:
   (a) The pressure is tripled while the absolute temperature is doubled.
   (b) The absolute temperature is doubled while the pressure is reduced to half.
   (c) The pressure and the absolute temperature are both doubled.
5. What is the major difference between a scientific law and a theory? Use Charles's law and the kinetic molecular theory to support your answer.
6. Convert each of the following gas pressures to units of kilopascals:
   (a) A mercury barometer gives the atmospheric pressure as 745 mm Hg.
   (b) A vacuum pump reduces the pressure in a container to 150 Pa.
   (c) An industrial process maintains a pressure of 2.50 atm in the reaction chamber.
7. Convert each of the following temperatures into kelvin temperatures:
   (a) freezing point of water
   (b) 21°C room temperature
   (c) 37°C body temperature
   (d) absolute zero
8. Pressurized hydrogen gas is used to fuel some prototype automobiles. What is the new volume of a 28.8-L sample of hydrogen for which the pressure is increased from 100 kPa to 350 kPa? State the assumptions that you need to make in order to answer this question.
9. One of the most common uses of carbon dioxide gas is carbonating beverages such as soft drinks.
   (a) Squeezing a plastic bottle increases pressure inside the bottle. What is the new volume of a 300-mL sample of carbon dioxide gas when the pressure doubles?
   (b) A 2-L bottle of pop, containing 300 mL of carbon dioxide gas at 125 kPa, is removed from the refrigerator. The temperature of the bottle and its contents increases from 7°C to 30°C. What is the new pressure of the carbon dioxide gas?
   (c) Explain why a can of carbonated pop sometimes overflows when opened.

10 A glass container can hold an internal pressure of only 195 kPa before breaking. The container is filled with a gas at 19.5°C and 96.7 kPa and then heated. Predict the temperature at which the container will break.

11. Electrical power plants and ships commonly use steam to drive turbines, producing mechanical energy from the pressure of the steam. The rotating turbine is connected to a generator that produces electricity. Steam enters a turbine at a high temperature and pressure and exits, still a gas, at a lower temperature and pressure.

Determine the final pressure of steam that is converted from 10.0 kL at 600 kPa and 150°C to 18.0 kL at 110°C.

12. Weather folklore tells of many signs that warn of approaching storms, such as wind rushing out of caves and vegetation floating to the surface of ponds. These phenomena are most likely due to the drop in atmospheric pressure associated with an advancing low-pressure area, which often accompanies a storm.
    (a) A cave holds 3000 m^3 of air when the atmospheric pressure is 103 kPa. Calculate the increase in volume of the air when the atmospheric pressure drops to 97 kPa.
    (b) Decaying vegetation at the bottom of a pond contains thousands of bubbles of trapped methane gas. If one piece of vegetation contains a total volume of 100.0 mL of methane gas at 103 kPa, calculate the volume of gas when the pressure drops to 97 kPa.
    (c) In light of your answers to (a) and (b), explain why wind rushing out of caves and decaying vegetation floating to the surface of ponds might be evidence of an advancing storm.
13. Many campers use propane as a fuel for cooking. If a tank contains 4.54 kg of propane, what volume of propane gas could be supplied at 12°C and 96.5 kPa?
14. An advertiser needs 5000 helium-filled balloons for a special promotion. If each balloon has a volume of 7.5 L on a day when the temperature is 18°C and the atmospheric pressure is 102.7 kPa, calculate the mass of helium required.
15. What amount of air (in moles) is present in an empty room with the dimensions 2.95 m × 3.50 m × 2.45 m at 21°C and 99.5 kPa?

Applying Inquiry Skills

16. For the purpose of testing Boyle's law, a student performs an experiment similar to Investigation 9.2.1, which studied the relationship between the volume and pressure of a gas. Complete the **Hypothesis**, **Analysis**, and **Evaluation** sections of the following lab report.

    Question
    What is the relationship between the volume and pressure of a gas?

    Hypothesis
    (a) Answer the Question and provide your reasoning.

    Experimental Design
    A syringe is sealed at the tip and then placed inside a 2-L plastic bottle. The pressure on the syringe is gradually increased by pumping air into the bottle. Gas pressure in the bottle is the independent variable, and gas volume in the syringe is the dependent variable.

    Evidence

    Air Pressure and Volume in a Syringe

Pressure (kPa)	Volume (mL)
100	50
150	33
200	25
250	20
300	17

    Analysis
    (b) What is the answer to the Question?

    Evaluation
    (c) Evaluate your Hypothesis. Is Boyle's law supported by the Evidence gathered in this investigation?
17. Design an experiment using a syringe that is sealed at the tip to determine the value of absolute zero.

Making Connections

18. Describe two natural phenomena and two technological products or processes using the gas laws.
19. Propane tanks for barbecues have a limited life. Propane refill stations can refuse to fill a tank and consumers may be stuck with the tank because recycling and disposal sites often do not accept old tanks. Suggest some reasons for the problems with refilling and disposal.
20. Describe briefly how air bags work and identify some risks and benefits of their use.
21. Air has many components.
    (a) List some categories that can be used to classify the components of air. Provide some examples for each category of gas.
    (b) Choose one of the components of air and tell its story from a variety of perspectives, for example, from scientific, technological, economic, and environmental perspectives.

Exploring

22. Use the Internet to research one of the noble gases to determine its abundance in the atmosphere, its commercial uses, and how commercial quantities are obtained.

    Follow the links for Nelson Chemistry 11, Chapter 9 Review.

    GO TO www.science.nelson.com