Chapter 11
The Behaviour of Gases

How many gases are shown in the main photograph on the opposite page? The only gas you can actually *see* is the water vapour in the clouds. In fact, however, many gases are present. Our atmosphere is made up of a mixture of different gases. The most important of these gases is oxygen, which we need to breathe. But did you know that most of the air you breathe is composed of nitrogen gas? This fact is well known to deep-sea divers, who encounter problems with nitrogen gas when diving far below the surface of the ocean.

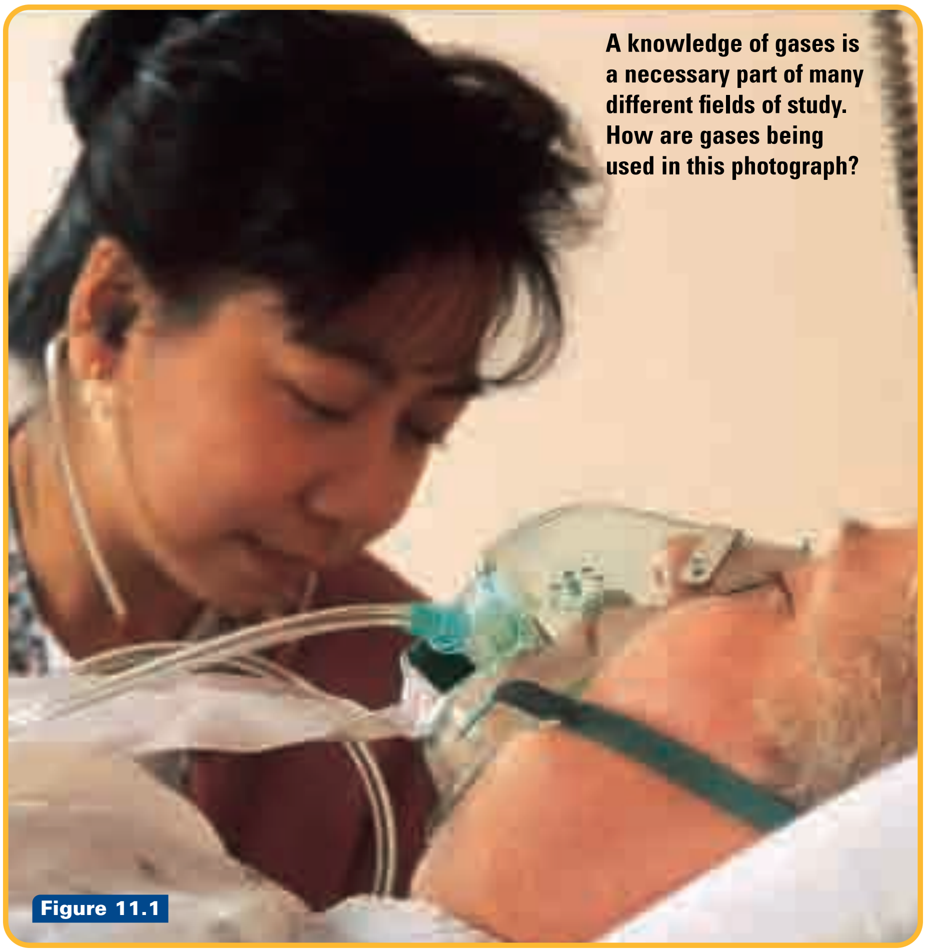
Gases are important in many different areas, from medical technology to the food industry. In this chapter, you will learn how particles behave in the gaseous, liquid, and solid states. You will also learn about laws that predict the behaviour of gases under different conditions of pressure, temperature, and volume. As well, you will discover some of the important ways in which gases are used in everyday life.

Chapter Preview

Concepts and Skills You Will Need

Before you begin this chapter, review the following concepts and skills:

- significant digits (Chapter 1, Section 1.2)
- unit analysis problem solving (Appendix D)
- the behaviour of atoms and molecules (Chapter 3, Sections 3.1, 3.2, and 3.3)



A knowledge of gases is a necessary part of many different fields of study. How are gases being used in this photograph?

Figure 11.1

11.1 States of Matter and the Kinetic Molecular Theory

Section Preview/ Specific Expectations

In this section, you will

- **explain** states of matter in terms of intermolecular forces and the motion of particles
- **perform** an ExpressLab to determine if gases occupy space
- **describe** a gas using the kinetic molecular theory
- **communicate** your understanding of the following terms and concepts: *condensation, kinetic molecular theory of gases, ideal gas*

Most of the universe is composed of plasma, a state of matter that exists at incredibly high temperatures (>5000˚C). Under normal conditions, matter on Earth can only exist in the other three physical states, namely, the solid, liquid, or gaseous states. As you learned in an earlier course, the particle theory describes matter in all states as being composed of tiny invisible particles, which can be atoms, ions, or molecules. In this section, you will learn how these particles behave in each state. You will also learn about the forces that cause their behaviour.

Solids and Liquids

In previous courses, you learned about the properties of the different states of matter. You may recall that both solids and liquids are incompressible. That is, the particles cannot squeeze closer together, or compress. The incompressible nature of solids and liquids is not due to the fact that particles are touching. On the contrary, the particle theory states that there is empty space between all particles of matter. The incompressibility of solids and liquids arises instead from the fact that these particles cannot move independently of each other. That is, the movement of one particle affects the movement of other particles, or is restricted by them.

This is especially true for solids. The particles of a solid are held together in a framework, called a crystal lattice. In a crystal lattice, the positions of solid particles are relatively fixed. This explains why solids have definite shapes: the particles are unable to slip past each other and thus change the shape of the solid.

Like solids, liquid particles cannot move independently of one another. They can slip past each other enough, however, to flow and change shape.

Another property of states of matter is their motion. According to the particle theory, all particles that make up matter are in constant motion. In solids, the range of motion of the particles is the most restricted. Each particle of a solid is only able to vibrate around a fixed point in the lattice (see Figure 11.2). This is called *vibrational motion*. Since solid particles are fixed in space, the degree of disorder is very low.

Figure 11.2 **The arrangement of particles in a solid. Particles vibrate in a fixed position relative to one another. They are unable to move past each other.**

Particles in the liquid state can move more freely than particles in the solid state, although not entirely independently. Liquid particles move with *rotational motion* as well as vibrational motion. This means that the particles can rotate and change position. It explains why liquids are able to flow and change shape, but keep the same volume. Since liquid particles move around more than solid particles, the liquid state has a higher degree of disorder, as shown in Figure 11.3.

To summarise the discussion, particles in solids and liquids are incompressible and thus have definite volumes. The particles in each state cannot move independently of each other. Therefore they are relatively restricted in their motion. How are the properties of gases different from those of solids and liquids?

The Gas State

Unlike solids and liquids, particles in the gaseous state are able to move independently of one another. Gas particles are able to move from one point in space to another. This is called *translational motion*. Thus, gas particles move with all three types of motion: vibrational, rotational, and translational. Gas particles move through space in random fashion. However, they do travel in straight lines until their course is altered by collisions with other particles. Because gas particles move freely, there is a high degree of disorder in a gaseous state.

Gas particles move much faster than liquid particles. Liquids always flow to the lowest point because they are still greatly influenced by gravity. Because gas particles move so quickly, gravity does not affect them as much. Gases flow in all directions, including upward against gravity, until all of the available empty space is occupied. This is why gases expand to fill a container. (See Figure 11.4.)

Gases can be compressed, unlike both solids and liquids. What is different about their particle arrangement that allows for this? The space between gas particles is much larger than the space between liquid or solid particles. Even if gas particles are moved closer together through compression, the distance between each particle is still very large. The particles remain in the gaseous state. When gas molecules are compressed further, eventually the forces between molecules become strong enough to hold the gas molecules together. At this point, the gas changes to the liquid state. This is known as **condensation**.

Figure 11.3 Particles in liquids are not held in a fixed position relative to other particles. They can slide over and past one another.

Forces Between Particles

You have examined some properties of solids, liquids, and gases. You have seen how the motion of the particles affects these properties. Now you will examine how the particles affect each other.

The particle theory states that there are attractive forces between particles. The weaker the attractive force is between particles, the freer the particles are to move. Therefore, attractive forces between particles are at their strongest in the solid state. Attractive forces are at their weakest in the gaseous state.

The strength of attractive forces between particles in any physical state depends on two major factors: type of force and temperature. The effect of temperature, or kinetic energy, on the state of a substance will be covered in greater detail later on in this section.

Attractions Between Charged Particles

What types of attractive forces exist between particles? In Chapter 3, you learned that oppositely charged particles attract each other due to *electrostatic attraction*. Ionic bonding is one example of electrostatic attraction. A positive ion (an atom or molecule that has lost electrons) is attracted to a negative ion (an atom or molecule that has gained electrons). Ions form very strong *ionic bonds*. Since these attractive forces are so strong, ionic compounds usually exist in nature as solids. For example, table salt (sodium chloride, NaCl) is a solid crystalline substance. It has a high melting point and a high boiling point.

Figure 11.4 Gas particles move freely in all directions, bouncing off each other as well as off the walls of their container.

CHECKPOINT

What is a dipole? Go back to Chapter 3 or Chapter 8 to refresh your memory.

Attractions Between Polar Molecules

Not all particles are charged, but attractions can still form between them. You learned about *intermolecular forces* in Chapter 3. Intermolecular forces are forces that exist between neutral molecules, or between molecules and ions.

You know that some molecules are polar due to their asymmetrical shapes. Sulfur dioxide (SO_2) is one example of a polar molecule. These molecules have a permanent dipole effect. This means that one end of the molecule is more positive, and the other end is more negative.

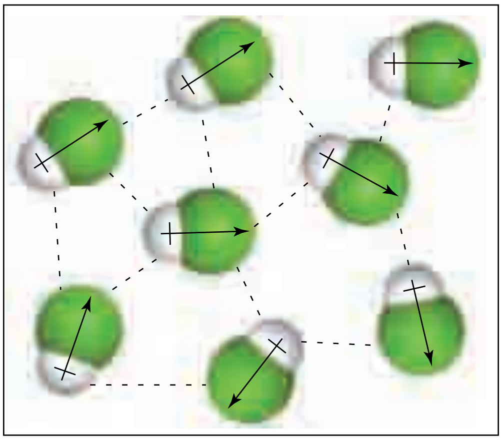


Figure 11.5 **Dipole–dipole interactions between polar HCl molecules**

Polar molecules attract ions and other polar molecules. The partially positive end of one molecule is attracted to the partially negative end of another molecule. This pattern continues throughout the substance. These *dipole–dipole* forces of attraction are not as strong as ionic bonds. Thus substances made up of polar molecules can exist as liquids and gases. For example, ethanol (CH_3CH_2OH) is polar, and exists as a liquid. You know this liquid as rubbing alcohol. Hydrogen chloride (HCl) is also polar. It is a gas under normal conditions as shown in Figure 11.5.

Attractions Between Non-Polar Molecules

What about substances made of non-polar molecules? You learned in Chapter 3 that *weak dispersion forces* form between non-polar molecules. As temporary dipoles form, they cause molecules to move closer together. However, these attractions are temporary and weak. Thus, most small non-polar molecules do not hold together long enough to maintain their solid or liquid forms. As a result, most small non-polar molecules exist as gases at room temperature. For example, carbon dioxide (CO_2) is a gas at normal temperatures.

CHECKPOINT

Hydrogen bonding is a very strong type of dipole-dipole interaction. Go back to Chapter 8 to review how this type of intermolecular force works.

The Relationship Between Size and State

Dispersion forces are also the primary forces of attraction between large non-polar molecules. However, as these molecules increase in size, their melting and boiling points rise. For example, methane (CH_4) is a small non-polar molecule. It has a very low boiling point and exists as a gas at room temperature. Pentane (C_5H_{12}) is a larger non-polar molecule. It has a higher boiling point, so it exists as a liquid at room temperature. Pentane has more sites along its length than methane does where temporary dipoles can form. The dispersion forces add up, so that it takes more energy overall to separate the molecules. This leads to a higher boiling point.

Electronic Learning Partner

If you are having difficulty visualizing molecules in the different states of matter, go to the Chemistry 11 Electronic Learning Partner.

To summarize, the state of a substance depends on the forces between the particles of that substance. If the forces are very strong, that substance is likely to exist as a solid. If the forces are weaker, that substance will exist as a liquid, or as a gas. The state of a non-polar substance also depends on the size of the molecule. Smaller non-polar molecules are more likely to be gases. Larger non-polar molecules will probably exist as liquids or even solids. Table 11.1 shows the forces discussed in this section ranked in order of strength.

Table 11.1 Attractive Forces

strong forces ← **weak forces**

Force	Ionic	Polar (dipole–dipole)	Dispersion
Type of force	between ions (intramolecular)	between molecules (intermolecular)	between molecules (intermolecular)
State	usually solid	liquid or gas (can also be solid)	liquid or gas
Example	$NaCl_{(s)}$	$CH_3CH_2OH_{(\ell)}$, $HCl_{(g)}$	$C_5H_{12(\ell)}$, $CH_{4(g)}$, $CO_{2(g)}$

The Effect of Kinetic Energy on the State of a Substance

There is one more factor that affects the state of a substance: temperature, which is related to kinetic energy. A hotter substance with high kinetic energy is more likely to overcome attractive forces between molecules, and exist as a gas. A cooler substance with low kinetic energy is more likely to be a solid or a liquid. This explains why heating a substance causes a change in state. When a solid is heated, it gains kinetic energy. Eventually it will melt, and become a liquid. When you add kinetic energy by heating a liquid, it will boil and become a gas. Earlier in this section, you learned that gases move much more quickly than liquids or solids. This is because gases have high kinetic energy.

CHECKPOINT

What kind of molecular forces would you expect for KI, SCl_6, and SiO_2? Use diagrams to explain your answer. Compare their melting and boiling points.

Kinetic Molecular Theory of Gases

The particle theory of matter does not discuss the kinetic energy of particles. Kinetic energy is important, however, when describing the unique properties of gases.

The **kinetic molecular theory of gases** makes the following assumptions:

- The volume of an individual gas molecule is negligible compared to the volume of the container holding the gas. This means that individual gas molecules, with virtually no volume of their own, are extremely far apart and most of the container is "empty" space.
- There are neither attractive nor repulsive forces between gas molecules,
- Gas molecules have high translational energy. They move randomly in all directions, in straight lines. (See Figure 11.6, on page 423.)
- When gas molecules collide with each other or with a container wall, the collisions are perfectly elastic. This means that when gas molecules collide, somewhat like billiard balls, there is no loss of kinetic energy.
- The average kinetic energy of gas molecules is directly related to the temperature. The greater the temperature, the greater the average motion of the molecules and the greater their average kinetic energy.

The kinetic molecular theory describes a hypothetical gas called an **ideal gas**. In an ideal gas, the gas particles take up hardly any space. Also, the particles of an ideal gas do not attract each other.

ExpressLab Do Gases Take Up Space?

Not everything can be seen with the unaided eye. Looking at a solid or a liquid, it is easy to see that they have mass and volume. Most gases are colourless. How can we "see" the volume of a gas?

Materials

1 L or 2 L clean plastic soft drink or juice bottle
round balloon
pointed scissors

Safety Precautions

Be careful with the sharp point of the scissors when piercing the plastic bottle.

Procedure

1. Insert the balloon into the bottle, holding the open end. Stretch the open end of the balloon over the lip of the bottle.
2. Step 3 will ask you to inflate the balloon as large as you can. Before you do Step 3, predict how much you will be able to inflate the balloon. Record your prediction in your notebook.
3. Inflate the balloon inside the bottle. How large did it get? Record your observations in your notebook.
4. Using the sharp end of a pair of scissors, puncture a hole in the middle of the bottom of the plastic bottle. Inflate the balloon again. Record your observations in your notebook.

Analysis

1. Was your prediction in Step 2 verified in Step 3? If you had problems inflating the balloon in Step 3, explain why.
2. Was there a difference in how much you were able to inflate the balloon after you punctured a hole in the bottle? If there was, explain why.
3. From your observations in this activity, do gases take up space? Explain your answer.

Why Use the Kinetic Molecular Theory?

How and why did scientists formulate the kinetic molecular theory? Experiments into gas behaviour demonstrate that, under normal temperatures and pressures, nearly all gases behave in similar and predictable ways. The properties and behaviours of real gases can be generalized into a theory of an ideal gas. This generalization makes it possible for us to calculate mathematically, with a high degree of accuracy, how real gases will behave under varying conditions.

Of course, no gas is really "ideal." The ideal gas theory ignores certain facts about real gases. For example, an ideal gas particle does not take up any space. In fact, you know that all particles of matter must take up space. Gas particles are small and far apart, however. Thus the space occupied by the particles is insignificant compared to the total volume of the container. You will learn more about the behaviour of real gases in Chapter 12.

Figure 11.6 This diagram shows the possible path of one gas molecule inside a volleyball. In a sample of gas, there are countless molecules moving in straight lines. They rebound off each other and the inner wall of the volleyball.

Section Wrap-up

The molecular-level interpretation of gas behaviour given by the kinetic molecular theory helps to explain the macroscopic, or "larger picture," properties of gases in the real world. One of the most important properties of gases is their compressibility—how they react to the application of an external force. In the next section, you will observe how gases behave under pressure. Later in this chapter, you will learn about some interesting applications of pressurized gases.

CHEM FACT

Oxygen molecules in the atmosphere, at room temperature, travel at an average speed of 443 m/s. This is approximately 1600 km/h!

Section Review

1. K/U Using the kinetic molecular theory of matter, explain each of the following observations.
   - **(a)** Gases are more compressible than liquids.
   - **(b)** The density of gases is less than that of solids.
2. K/U In your own words, describe the characteristics of an ideal gas.
3. MC Using your knowledge of intermolecular forces, predict the state of each substance at room temperature. Explain your answer.
   - **(a)** hexane (C_6H_{14})
   - **(b)** hydrogen fluoride (HF)
   - **(c)** potassium chloride (KCl)
4. I Explain each of the following observations.
   - **(a)** Metals expand when heated, yet contract in cold weather.
   - **(b)** Gases have no fixed volume.
   - **(c)** A certain amount of moles of water occupies much more space as a gas than as a liquid.
5. C How does the degree of disorder of a gas compare to that of a liquid or a solid? Explain your answer.
6. K/U Describe the motion of a gas particle.
7. K/U What effect does heating have on the particles of a liquid?
8. **(a)** C Draw five boxes in your notebook. Inside them, illustrate the motion of gas particles according to the kinetic molecular theory.
   - **(b)** Draw another five boxes underneath the five boxes in (a). Illustrate how you think the molecules of a real gas might move in comparison.

11.2 Gas Pressure and Volume

Section Preview/ Specific Expectations

In this section, you will

- **perform** experiments to determine the quantitative and graphical relationships between pressure and volume in an ideal gas
- **solve** problems using Boyle's law
- **review** your understanding of the following terms and concepts: *newton, pascal, kilopascal, pressure, volume*
- **interconvert** units of pressure
- **communicate** your understanding of the following terms: *closed system, pressure, pascal, kilopascals, mm Hg, torr, atmospheres, standard atmospheric pressure, Boyle's law*

The earliest use of pressure in English referred to a burden or worry troubling a person's mind. Scientists found this a useful mental model to picture what happens when force is applied to a specific area. They adopted the word pressure to describe any application of force over an area.

Throughout the rest of this chapter, you will discover how gases behave when they are under pressure in a closed system. A **closed system** is one with a constant amount of moles of a substance. It is not open to the atmosphere. Gases in closed systems, from CO_2 in fire extinguishers to O_2 in oxygen tanks, perform important functions in our lives. Understanding the behaviour of gases in closed systems is essential to our safe and effective use of gases.

How is Pressure Calculated?

As you learned in previous studies, **pressure** is defined in physical terms as the force exerted on an object per unit of surface area ($P = F/A$). One commonly used SI unit of pressure is the **pascal** (**Pa**), equal to 1 N/m^2. More often, pressure is reported in **kilopascals** (**kPa**), equal to 1000 Pa. (You will learn about other units of pressure later in this section.)

Assume a student with a mass of 51.0 kg is sitting on a chair. The force the student applies to the chair is 500.0 N. If the surface area of the chair seat is 0.05 m^2, the pressure the student exerts is

$$P = \frac{F}{A} = \frac{500.0\text{ N}}{0.05\text{ m}^2} = 10\,000\text{ N/m}^2 = 10\,000\text{ Pa} = 10.0\text{ kPa}$$

Figure 11.7 shows how a decrease in surface area can dramatically increase pressure.

Figure 11.7 **A woman with a mass of 50.0 kg exerts a pressure of about 21 kPa on the floor as she walks. If another woman with an equal mass is wearing high heels, she will exert a pressure of about 5000 kPa as her heel hits the floor. This pressure is approximately 240 times greater than if she were wearing flat shoes!**

How does a gas exert pressure? In a sense, it cannot exert measurable pressure in the same way that a solid or liquid can. The pressure of a gas is determined by the kinetic motion of its component molecules. Suppose hundreds of billions of gas molecules are in random motion, striking the entire inner surface of their container. Each collision exerts a force on the container's inner surface.

Picture inflating a basketball. As you add more and more air to it, more molecules collide against the inside wall of the basketball. Each collision exerts a force on the basketball's inner surface area. The collective number of collisions as well as the strength of the force form the net or overall gas pressure. Since the molecules move in all directions, the net pressure exerted will be equal throughout. (Figure 11.8 illustrates this.)

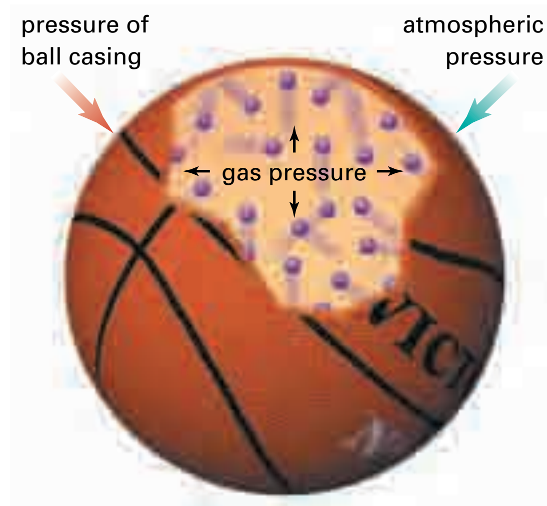


Figure 11.8 This diagram shows gas particles exerting pressure as they bounce off the inner surface of a basketball

Atmospheric Pressure

Despite the popular expression, you can't carry the world on your shoulders! Scientists estimate that the lithosphere, or solid Earth, has a mass of 6.0×10^{24} kg. The hydrosphere, or the portion of Earth covered by water, has an estimated mass of 1.4×10^{21} kg.

Figure 11.9 The column of air above one square metre (1 m^2) at sea level and 0°C exerts a pressure of 101 325 Pa (1 Pa = 1 N/m^2). This is equivalent to a mass of about 10000 kg over an area of 1 m^2!

You do constantly experience the pressure exerted by Earth's atmosphere. Scientists estimate that the atmosphere has a mass of 5.1×10^{18} kg. Thus air molecules, which have mass, are being pulled down by gravity and are exerting pressure on all objects on Earth. Figure 11.9 shows how much pressure is exerted by the atmosphere over an area of 1 m^2.

COURSE CHALLENGE

What are the main factors determining atmospheric pressure on Earth? In your Chemistry Course Challenge, you will consider the atmosphere on a newly discovered planet. Why might the atmospheric pressure on another planet be different?

mind STRETCH

Using the Internet or another source, find the average atmospheric pressure at the top of Mount Everest. If it takes four minutes to boil a soft-boiled egg at sea level, how long would it take to boil an egg at the top of Mount Everest?

Early Studies of Atmospheric Pressure

In the early seventeenth century, the Italian scientist Galileo Galilei (1564–1642) developed a suction pump. It used air to lift water up to the surface from about 10 m underground. When drawn from greater depths, the column of water collapsed before it reached ground level. Galileo concluded that the water could not be pumped higher because it had reached the "limit of vacuum." Pumping from any depth beyond 10 m required a greater pressure from the suction pump than was provided by the atmosphere. However, Galileo did not know exactly how water was being moved up the tube.

From 1641 to 1642, Evangelista Torricelli (1608–1647) served as Galileo's secretary. He continued Galileo's experiments and concluded that the weight of air was pushing down on the rest of the water. The weight of the air pushed water up the column. This was a logical conclusion since gas molecules, like all matter on Earth, are pulled down by the force of gravity.

Torricelli did further calculations involving the weight of the atmosphere pressing on the water. He then improved upon the experiment by using mercury, which has a density 13.6 times greater than water. He designed the apparatus which we now know as the barometer, shown in Figure 11.10. Torricelli filled a glass tube of 1 cm diameter, closed at one end, with mercury. He then inverted the tube in a dish of mercury. Some of the mercury ran out of the tube. But about 760 mm of mercury remained in the tube. This is about 13.6 times less than the height of water that Galileo could pump. It is the air pushing on the mercury in the dish that keeps this 760 mm of mercury in the tube.

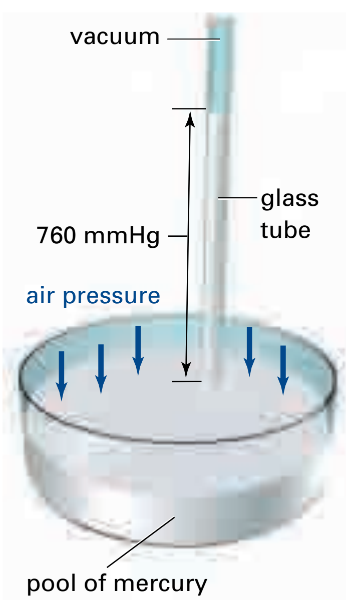


Figure 11.10 Torricelli's barometer

Changes in Atmospheric Pressure

At first, Torricelli considered his experiment a failure. The height of the mercury column did not remain constant at 760 mm, but changed slightly as the weather and air temperature changed. As you learned in Grade 10, these small changes in the height of the mercury column provide us with valuable information. We can predict the weather, in part, by looking at changes in atmospheric pressure. For example, if the atmospheric pressure decreases suddenly, a storm may be on the way.

Atmospheric pressure affects us in other ways, too. People who live at high elevations, such as in the Rocky Mountains, have less mass of air above them. At lower atmospheric pressure, the boiling point of water decreases. Because the water boils at a lower temperature, it takes longer to cook food in boiling water on a mountain than at sea level.

The following ExpressLab demonstrates atmospheric pressure in a dramatic way, showing you firsthand the tremendous pressure that the air around us can exert.

ExpressLab — Observing Atmospheric Pressure in Action

In this ExpressLab, you will see just how powerful atmospheric pressure can be.

Materials

empty, clean soft drink can
hot plate
beaker tongs
large beaker of ice water
10 mL graduated cylinder
5 mL of water

Safety Precautions

Use safety goggles during this activity. Handle the heated can carefully using the beaker tongs.

Procedure

1. Pour 5 mL of water into the soft drink can.
2. Heat the can on the hot plate until steam begins rising from the opening of the can.
3. Using the beaker tongs, quickly invert the can into the large beaker of ice water so that the opening of the can is just under the surface of the water. Observe carefully.

Analysis

1. What happened to the air and water molecules inside the can when it was heated?
2. Explain what happened to the can when it was placed in the ice water.

Extension

3. Calculate the surface area of the outside of the can exposed to the atmosphere. Assuming an atmospheric pressure of 100.0 kPa, how much force was applied to the can by the atmosphere?

Tools & Techniques

High Pressure Injectors

When gas molecules under high pressure are allowed to escape and expand, they release kinetic energy. This energy has been harnessed to benefit human health by powering high-pressure injectors, better known as jet injectors.

Jet injectors are hypodermic syringes that use high-pressure gas instead of a needle to inject vaccines under a patient's skin. Dr. Robert Higson (1913–1996) developed the "Hypospray" in the 1940s. This device used spring pressure against a plunger to force a vaccine through a tiny nozzle at 1000 km/h. The pressure was enough to drive the vaccine into the tissue of a patient's arm without breaking the skin.

Today jet injectors use a tank of compressed gas and an automatic vaccine dispenser that works through a pistol-like injector. When triggered, the new device releases a measured dose of vaccine into a sterile chamber and a small volume of gas through a hose. As the compressed gas expands, it forces the vaccine at high velocity through the injector's nozzle. The process is fast and simple, ideal for performing mass vaccinations. It also eliminates the problems of disposing of used syringes. Also, it protects medical personnel from infection through accidental needle pricks.

In 1958, Higson led a team that inoculated about 90 000 people in Asia and Africa against polio, typhoid, and cholera using the Hypospray. In 1965, the United Nation's World Health Organization used the jet injector. The organization freeze-dried vaccines for its successful worldwide smallpox eradication program. For its key role in eliminating smallpox from the list of human diseases, the jet injector earned a new name—the "Peace Pistol."

Math LINK

You may encounter tire pressure gauges that are calibrated in pounds per square inch (psi). The conversion factor between kPa and psi is 101.3 kPa = 14.7 psi. To convert a tire pressure of 28.0 psi:

$$\frac{x\ \text{KPa}}{28.0\ \cancel{\text{psi}}} = \frac{101.3\ \text{kPa}}{14.7\ \cancel{\text{psi}}}$$

$$x = 28.0\ \text{psi} \times \frac{101.3\ \text{kPa}}{14.7\ \text{psi}}$$

$$= 193\ \text{kPa}$$

What would a tire pressure of 27.3 psi be in kPa? What would the pressure of 198.7 kPa be in psi?

Units of Pressure

For many years, atmospheric pressure was measured in millimetres of mercury (**mm Hg**). In the British Commonwealth and the United States, inches of mercury were used. Standard atmospheric pressure, the pressure of the atmosphere at sea level and 0˚C, is 760 mm Hg. More recently, in honour of the work of Torricelli, standard atmospheric pressure has been defined as 760 torr. **1 torr** represents a column of mercury 1 mm in height at 0˚C. Another common unit for measuring pressure is **atmospheres** (**atm**), where 1 atm is equivalent to 760 torr. While mm Hg, torr, and atm are still used to measure pressure, especially in technological and medical applications, the SI units are pascals (Pa) or kilopascals (kPa).

In other words, **standard atmospheric pressure** at 0˚C is equivalent to:

$$760\ \text{mm Hg} = 760\ \text{torr} = 1\ \text{atm} = 101.3\ \text{kPa}$$

Using this relationship, we can convert from one unit to another. For example, a pressure of 100.0 kPa is equivalent to

$$100.0\ \cancel{\text{kPa}} \times \frac{760.0\ \text{torr}}{101.3\ \cancel{\text{kPa}}} = 750.2\ \text{torr}$$

The Relationship Between Pressure and Volume

Figure 11.11 shows a meteorologist preparing to release a weather balloon partially filled with helium gas. As the balloon rises, atmospheric pressure decreases. The volume of the balloon increases.

Figure 11.11 **Weather balloons are partially inflated with helium. They carry specialised instruments to measure varying atmospheric conditions such as pressure, temperature, and humidity.**

Since the helium atoms inside the balloon move randomly in all directions, they constantly bombard all the area inside the walls of the balloon, exerting a pressure. With decreasing atmospheric pressure, there are fewer air molecules to collide with the outside of the balloon. As the pressure outside the balloon becomes less than that inside the balloon, the balloon expands. Given an expandable container, such as a balloon, the volume occupied by the gas will increase when external pressure decreases. As external pressure *increases*, gas molecules are forced closer together. The volume of gas then *decreases*.

When we refer to the *volume of a gas*, we are in fact talking about the *volume of the container*. The definition of the volume of a gas is *the space available for gas molecules to move around in*. The kinetic molecular theory of gases assumes that the volume of each gas molecule is essentially zero. Thus, the amount of space for them to move around in is the volume of the container. For all gases, V_{gas} = the volume of the container holding the gas. (Do not confuse this with the *molar volume* of gases. You will learn about molar volume in Chapter 12.)

Think about how the relationship between pressure and area for solids would apply if you were testing a gas in a three-dimensional container. You know that according to the kinetic molecular theory, gas molecules exert pressure over the entire inside surface of their container. If the volume of the container is halved, what would happen to the pressure of the gas inside the container?

Robert Boyle (1627–1691) was an Irish scientist with an interest in chemistry. He investigated the relationship between pressure and volume of gases at constant temperatures. By making careful measurements of the volume of a trapped gas, he was able to describe what happened when the pressure exerted on the gas was increased. Figure 11.12 shows Boyle's experiment. Boyle measured the length of the column of trapped air compared to the length of the column of mercury. Since the length of the mercury column is directly related to its volume, Boyle was able to deduce the relationship between pressure and volume.

Electronic Learning Partner

Go to the Chemistry 11 Electronic Learning Partner for a demonsration of Boyle's law.

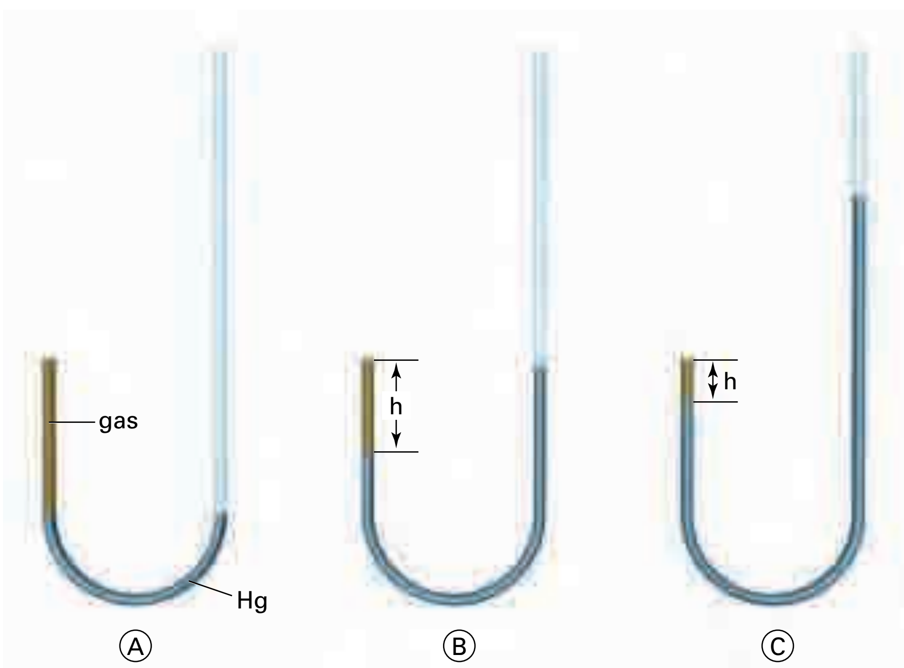


Figure 11.12 **When liquid mercury is added to the open tube, the pressure caused by the weight of mercury on the trapped gas increases. The volume of the trapped gas (b and c) decreases.**

Since mercury is a poisonous element, you will use a different method in Investigation 11-A to examine the relationship between the pressure and the volume of a gas.

Investigation 11-A

SKILL FOCUS

Performing and recording

Analyzing and interpreting

Identifying variables

The Relationship Between the Pressure and the Volume of a Gas

Boyle measured the variance in length of a column of trapped air. Since length is directly proportional to volume in a column with a regular diameter, this gave him an indirect measure of changes in volume of the air with increased pressure. In this activity, air is trapped inside a sealed plastic pipette. You can measure the volume of the trapped gas in terms of the length of the column of air, the way that Boyle did more than three hundred years ago. You will measure the applied pressure in terms of the number of turns of a clamp rather than in kPa.

Question

What is the relationship between the pressure and volume of a fixed amount of gas at a constant temperature?

Materials

thin stem plastic pipette with bulb
small C-clamp
metric ruler (with mm)
match or Bunsen burner
craft stick
coloured water

Safety Precautions

- Be very careful when sealing the end of the plastic pipette with a flame. The plastic will melt and may begin to burn. Hot, molten plastic can burn your skin.
- Do not inhale any of the fumes from the plastic.
- Before lighting the Bunsen burner, check that there are no flammable solvents nearby.

Procedure

1. Squeezing the bulb, draw enough coloured water into a pipette so that the water fills the bulb and extends about 2 to 3 cm down the stem. The rest of the stem should be filled with air.
2. Using a flame from a Bunsen burner, carefully seal the end of the pipette completely. Allow the pipette to cool for at least three minutes before completing the rest of the procedure.
3. Copy out the data table into your notes.

P (no. of turns)	V (mm)	1/V (mm)	P × V (turns · mm)
0			
1			

4. Break a craft stick in two. Place one half of the stick on either side of the pipette bulb. Tighten a small C-clamp around the bulb of the pipette so that the clamp just holds the bulb snugly (see the diagram).
5. Using a ruler, measure the length of trapped gas (the "volume") in millimetres. Record this in your data table.
6. Increase the "pressure" on the bulb by turning the handle of the clamp one half or one complete turn, depending on the size of your clamp. Record the "volume" of the trapped gas in your data table.
7. Repeat step 6 until you have made at least five complete turns.
8. Complete the data table. Calculate an average value for the $P \times V$ column.
9. Plot a graph of P (y-axis) versus V (x-axis).
10. Plot a graph of P (y-axis) versus $1/V$ (x-axis).

Analysis

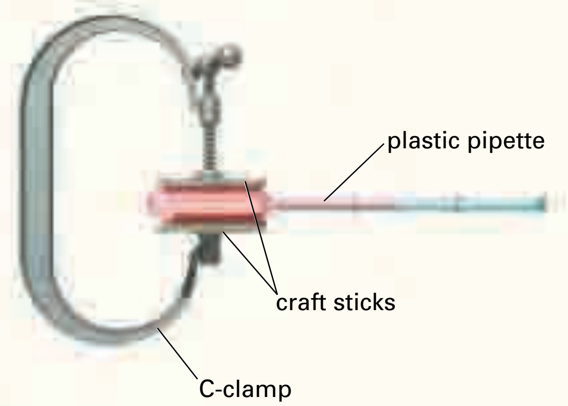
1. What relationship exists between volume and pressure, based on the data collected and the graphs produced?
2. Express this relationship mathematically. To help you to do this, look at the mathematical form of Boyle's law, located after this investigation.
3. Calculate the slope of the P vs. $1/V$ graph. How does this value compare with the average $P \times V$ value? Of what significance are these two values?
4. What changes in temperature occurred during the experiment? In the amount of trapped air? Explain how this may have affected your results.

Conclusion

In your own words, state the relationship between pressure and the volume of a gas.

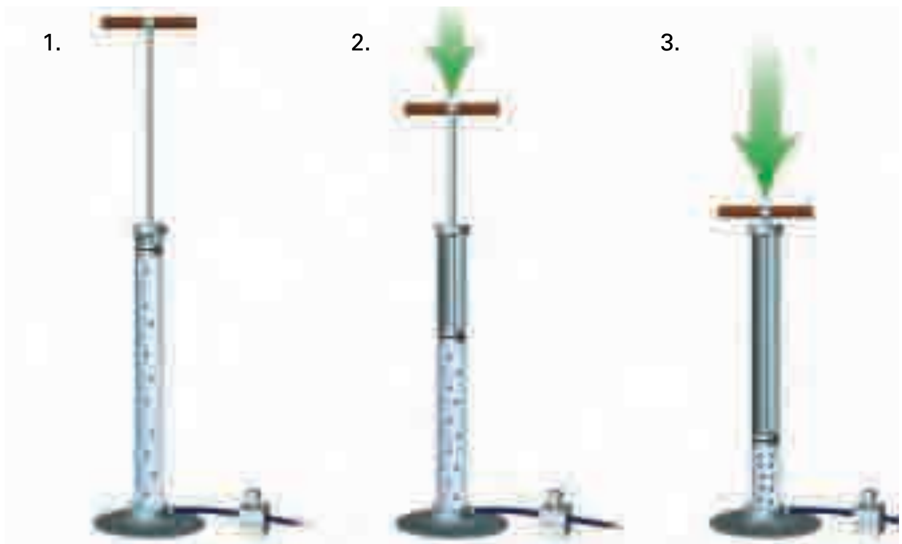
Extension

Using pressure probes and a graphing calculator or computer interface, investigate the relationship between the pressure and the volume of a gas. Produce a data table and graphical interpretation of these results.



Boyle's Law

In 1662, Robert Boyle stated that *the volume of a given amount of gas, at a constant temperature, varies inversely with the applied pressure*. In other words, as external pressure on a gas increases, the volume of the gas decreases by the same factor. This statement is known as **Boyle's law**. Figure 11.13 illustrates Boyle's law using a bicycle pump.



1. With the piston pulled all the way out, the air pressure inside a bicycle pump equals the air pressure outside.
2. If the piston is pushed halfway down the cylinder, the volume of the cylinder is decreased to one half its original value. The pressure doubles.
3. If the piston compresses the air to one fourth its original value, the pressure of the air inside the pump will be four times higher than its original value.

Figure 11.13

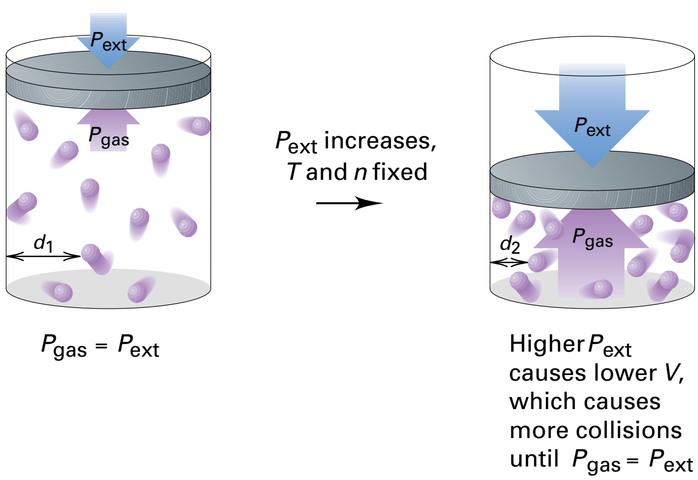


Figure 11.14 **At a given temperature, gas molecules travel an average distance (d_1) before they collide with the container wall. When the volume is decreased, the gas molecules travel a shorter distance (d_2) before striking the wall.**

Boyle found that this relationship held true for all gases as long as the temperature remained unchanged. As external pressure on a gas increases, the volume of the gas decreases. The gas molecules are forced closer together. However, if the volume of a gas decreases, then the gas molecules have to travel a shorter distance before they strike the container walls, as shown in Figure 11.14. Since they travel a shorter distance, gas molecules will strike the container walls more often per unit time. This increases the internal pressure of the gas. (With an increased volume, there are fewer collisions per unit time and a lower gas pressure is exerted.)

In other words, as the pressure of a closed system increases, its volume decreases. If the pressure is decreased by half, the volume doubles. We can write this relationship mathematically by using the proportionality symbol, α. V α $1/P$ means that volume is inversely proportional to the pressure.

Mathematically, the proportionality sign (α) can be removed by introducing a proportionality constant (k).

$$V \,\alpha\, \frac{1}{P}$$

$$V = \frac{1}{P} \times k \qquad \text{or} \qquad PV = k$$

In Investigation 11-A, when you plotted a graph of P versus $1/V$, you obtained a straight line. The slope of this line gives the value of the proportionality constant, k. If the pressure is tripled, the volume will decrease to one third of its original volume, such that $P \times V = k$. The value of k differs depending on the gas sample and the temperature. Remember, this mathematical relationship only applies if the temperature remains constant. A graph of P versus $1/V$ is shown in Figure 11.15.

For the gas sample at its initial conditions (i)

$$P_iV_i = k$$

If the gas sample is then subjected to a change in pressure, at its final conditions (f)

$$P_fV_f = k$$

Since the slope of the line (k) is constant, and since initial and final conditions are both equal to k, we can write

$$P_iV_i = P_fV_f$$

This mathematical relationship is another way of stating Boyle's law.

CHEM FACT

Scientific discoveries often happen simultaneously. Edmé Mariotte, a French scientist, investigated the pressure–volume relationship of gases independently of Boyle. He did not publish his work until 1676, fourteen years after Boyle had. In many European countries, the mathematical relationship between gas pressure and volume is known as Mariotte's Law.

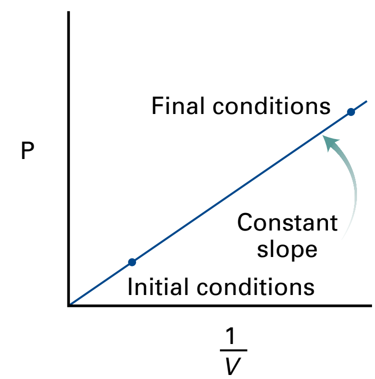


Figure 11.15 If a sample of gas at initial conditions has a change of pressure applied to it, its volume decreases proportionally, such that in its final state, $P \times V = k$.

Sample Problem

Boyle's Law: Calculating Volume

Problem

A sample of helium gas is collected at room temperature in a 4.50 L balloon at standard atmospheric pressure. The balloon is then submerged in a tub of water, also at room temperature, such that the external pressure is increased to 110.2 kPa. What will the final volume of the balloon become?

What Is Required?

You need to find the volume of the balloon after the pressure on the balloon has been increased. (V_f = ?)

What Is Given?

- You know the initial pressure and volume, and the final pressure.
  Initial pressure (P_i) = 101.3 kPa
  Initial volume (V_i) = 4.50 L
  Final pressure (P_f) = 110.2 kPa
- You know that the temperature does not change.

Plan Your Strategy

Algebraic method

- Since temperature is constant and pressure and volume have been given, you will need to use the Boyle's law formula.
- You can substitute numbers and units for the variables in the formula to solve for the unknown (V_f).

PROBLEM TIP

In this Sample Problem, you will see two different methods of solving the problem: the algebraic method and the ratio method. Choose the method you prefer to solve this type of problem.

Continued ...

Technology LINK

In 1998, a weather balloon carrying instruments to measure the ozone layer drifted off course. It veered into transatlantic air routes, where it posed a serious danger. By this time, the balloon had expanded in size to about the same volume as Toronto's Sky Dome. Two Canadian Air Force CF-18 jets directed over 1000 rounds of cannon fire at it, but could not bring the balloon down. It finally landed on an island off the coast of Finland. University of Toronto physicists have developed a new mechanism to prevent such an event from recurring. When an experiment has been completed, a parachute will return the scientific instruments to the ground. As the instruments fall, they will pull panels from the side of the balloon, causing it to plummet rapidly.

Can you think of other ways to solve the problem that faced the University of Toronto scientists? Create a set of blueprints for another technological solution, using CAD software if you have access to it.

Continued ...

FROM PAGE 433

Ratio method

- The pressure on the balloon increases. When this happens, if temperature remains the same, you know according to Boyle's law that the volume of the balloon will decrease.
- To find the final volume, you can multiply the initial volume of the balloon by a ratio of the two pressures that is less than one (i.e. $\frac{101.3 \text{ kPa}}{110.2 \text{ kPa}}$).

Act on Your Strategy

Algebraic method

$P_iV_i = P_fV_f$

$(101.3 \text{ kPa})(4.50 \text{ L}) = (110.2 \text{ kPa})(V_f)$

To isolate V_f, you need to divide both sides of the equation by 110.2 kPa.

$$\frac{(101.3 \cancel{\text{kPa}})(4.50 \text{ L})}{(110.2 \cancel{\text{kPa}})} = \frac{\cancel{(110.2 \text{ kPa})}(V_f)}{\cancel{(110.2 \text{ kPa})}}$$

$$\frac{(101.3 \cancel{\text{kPa}})(4.50 \text{ L})}{(110.2 \cancel{\text{kPa}})} = (V_f)$$

$$V_f = 4.137 \text{ L}$$

Ratio method

$$V_f = 4.50 \text{ L} \times \text{pressure ratio}$$
$$= 4.50 \text{ L} \times \frac{101.3 \cancel{\text{kPa}}}{110.2 \cancel{\text{kPa}}}$$
$$= 4.137 \text{ L}$$

Since the least number of significant digits in the question is three, the answer is:

$$V_f = 4.14 \text{ L}$$

Check Your Solution

- The units for the answer are in litres.
- When units cancel out, L remains.
- The volume of the balloon has decreased due to the increase in pressure.

Practice Problems

1. A 50.0 cm^3 sample of nitrogen gas is collected at 101.3 kPa. If the volume is reduced to 5.0 cm^3, and the temperature remains constant, what will the final pressure of the nitrogen be?

Continued ...

Continued ...
FROM PAGE 434

2. A weather balloon has a volume of 1000 L at a pressure of 740.0 torr. The balloon rises to a height of 1000 m where the atmospheric pressure is measured as 450.0 torr. Assuming there is no change in temperature, what is the final volume of the weather balloon?
3. A 45.0 cm^3 sample of nitrogen gas is collected at 1.0 atm. The nitrogen is compressed to a pressure of 10.0 atm. What is the final volume of the nitrogen if the temperature remains constant?
4. A 45.6 mL sample of gas at 490 torr is compressed to a certain volume at 3 atm. What is the new volume, in litres?

CHEM FACT

Now that you've finished practising Boyle's law problems, take a deep breath and relax. You have just illustrated Boyle's law! When you inhale, muscles in your torso expand your rib cage. The volume of your lungs increases. Since the pressure inside your lungs is decreased with the expansion in volume, outside air under higher pressure rushes in.

Section Wrap-up

In this section, you learned about the relationship between pressure and volume of a gas. This relationship is stated in Boyle's law. With knowledge of gas properties and behaviours, we are able to devise and improve upon technologies used everyday. You will learn about some of these important technologies later in this chapter. In the meantime, the next section examines how gases respond to changes in yet another variable: temperature.

Section Review

1. C Using the relationship 760 mm Hg = 760 torr = 1 atm = 101.3 kPa, convert each of the following units:
   - **(a)** 2.03 atm to kPa
   - **(b)** 85.2 kPa to atm
   - **(c)** 1.50 atm to torr
   - **(d)** 600 torr to kPa
2. K/U Use the kinetic molecular theory. Explain why the air pressure inside a capped syringe increases if the volume decreases from 15 cm^3 to 10 cm^3.
3. K/U Explain, using the kinetic molecular theory, why pressure is exerted by gases in all directions.
4. I A 1.00 L helium balloon is floating in the air on a day when the atmospheric pressure is 102.5 kPa and the temperature is 20.0°C. Suddenly, clouds appear and the pressure rapidly drops to 98.6 kPa at a temperature of 20.0°C. What is the new volume of the balloon?
5. I 0.750 L of oxygen gas is trapped at 101.3 kPa in a cylinder with a moveable piston. The piston is moved and the gas is compressed to a volume of 0.500 L. What is the final pressure applied to the oxygen gas if the temperature remains unchanged?
6. MC A student produces 38.3 mL of oxygen gas in a burette. The next day, there are 40.2 mL of gas in the burette at a pressure of 103 kPa. What was the pressure on the previous day, in torr? What might be happening to the weather in the student's neighbourhood?

11.3 Gases and Temperature Changes

Section Preview/ Specific Expectations

In this section, you will

- **perform** laboratory experiments investigating the effects of temperature changes on the volume of gases
- **solve** problems using Charles' law and Gay-Lussac's law
- **convert** units between the Celsius and Kelvin temperature scales
- **communicate** your understanding of the following terms: *Kelvin scale, absolute zero, Charles' law, Gay-Lussac's law, pressure-relief valve, fusible plugs*

CHECKPOINT

As you perform Investigation 11-B, keep in mind the ratio of 1/273 discovered by Charles. Can you recall from your previous studies a special significance for the number 273?

As you learned in Section 11.1, the average kinetic energy of gas molecules is directly related to the temperature. The greater the temperature, the greater the average motion of the molecules and the greater their average kinetic energy. In other words, the temperature of a substance is defined as the measure of the average kinetic energy of the molecules in that substance.

When substances are cooled, they lose kinetic energy. How does this affect their volume? Remember, for an ideal gas, we can think of its particles as having mass but no volume. When you perform Investigation 11-B on page 438, imagine air behaving as an ideal gas. How do you think air, in an expandable container will react to temperature changes?

Temperature and Volume

Sometimes, scientific discoveries are made well before any technological application of them can be envisioned. At other times, the desire to develop new technologies leads to experiments from which discoveries are made.

Jacques Charles (1746–1823), a French scientist, was the first to fill a balloon with hydrogen. He was also interested in hot-air balloons, which were being developed in France at the time. Charles investigated the expansion rates of nitrogen, oxygen, hydrogen, and carbon dioxide. He found that these gases all expanded by the same ratio. For each degree Celsius increase in temperature, all of these gases would expand by a certain fraction. This fraction was 1/273rd of their volume at 0°C. For each degree Celsius decrease in temperature, their volume would decrease by the same fraction. Thus, if a gas at 0°C were to be heated to 273°C, its volume would double. This held true when the pressure and the amount of gas remained constant. Figure 11.16 shows the expansion of the volume of gas in a hot-air balloon as the air inside the balloon is heated.

Figure 11.16 **These photographs show the gradual expansion of a hot air balloon. Since hot air is less dense than cooler air, the balloon rises.**

Chemistry Bulletin

Science Technology Society Environment

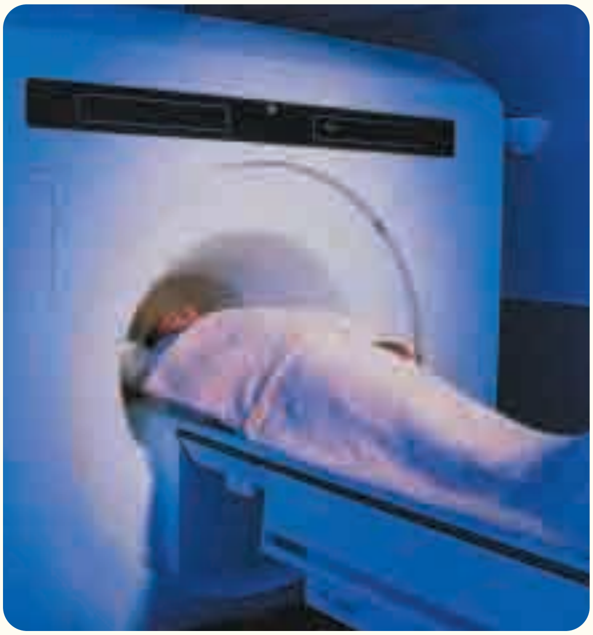
Gas Temperature and Cryogenics

Some people wish to be frozen when they die, trusting that future technologies will be able to revive them. This untested process is called *cryonics*. It was first suggested by Robert C.W. Ettinger in 1962. Ettinger has since set up his own Cryonics Institute in Michigan, where people can have their bodies frozen and stored. We still cannot freeze and re-animate higher animals, such as people. Cryogenic freezing, however, can suspend the life of tissues and organs used for transplants. Cryogenics has also made many other technologies possible. It has proven that science and science fiction are often closer than we think!

The term *cryogenics* can be applied to all temperatures below the normal boiling point of oxygen, or –183˚C.

When substances encounter temperatures this low, they often behave strangely. Liquid helium, for example, becomes a "superfluid" at temperatures below –270.97˚C. The principle of "superconductivity" is just as interesting Cryogenically cooled rings made out of metals such as lead and aluminum become "superconductors." They can keep currents travelling in circles for hours even after scientists have removed the original source of electricity.

Scientists have made use of the strange things that happen to matter at low temperatures. Superconductors have been used to make huge electromagnets, like the one at the Argonne National Laboratory near Chicago. Argonne's electromagnet can produce a magnetic field 134 000 times as strong as Earth's. It operates on relatively little power because of its superconducting capabilities. Such magnets are used in nuclear power research to find new nuclear particles. And while cryogenics has advanced nuclear science, it also helps scientists to study the effects of nuclear radiation. Scientists study cryogenically frozen atoms suspended in an irradiated state to understand how nuclear radiation can harm human health.



Cryogenically cooled magnets are used in MRI technology.

Cryogenics produces large-scale amounts of nitrogen and oxygen. Scuba divers and astronauts use compressed oxygen tanks that provide a six to eight hour supply of oxygen. Rocket engines use liquid oxygen as fuel. We use liquid nitrogen to make ammonia for fertilizers, to keep frozen foods cold during transport, and to fast-freeze these foods.

The list of applications of cryogenics is long and varied, and research continues. Maybe Ettinger is right and one day all of the occupants of the Cryonics Institute will live again!

Making Connections

1. What do you think might be some possible future applications of cryogenics?
2. Scientists use Dewar Flasks to contain cryogenic fluids. How do these flasks work? Do some research to find out.

SKILL FOCUS

Performing and recording

Analyzing and interpreting

Identifying variables

Investigation 11-B

The Relationship Between Temperature and Volume of a Gas

As you learned in Investigation 11-A, the length of a column of trapped air is directly proportional to its volume. In this investigation, you will see the effect of temperature changes on the volume of a gas, also measured in terms of the length of a column of trapped air.

Question

What is the relationship between the temperature and volume of a fixed amount of gas at a constant external pressure?

Materials

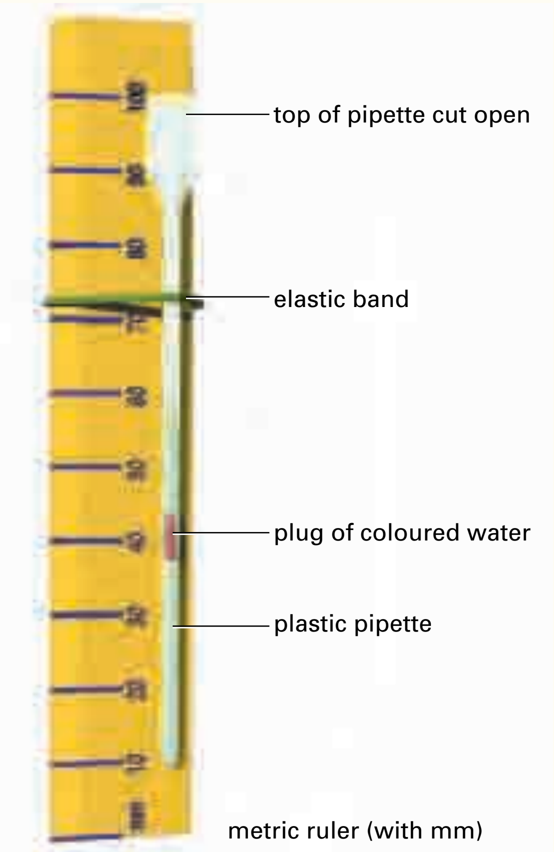
thin stem plastic pipette
metric ruler (with mm)
Celsius thermometer
400 mL beaker
ice
hot plate
match or Bunsen burner
2 elastics
scissors
coloured water
tap water

Safety Precautions

- Be very careful when sealing the end of the plastic pipette with a flame. The plastic will melt and may begin to burn. Hot, molten plastic can burn your skin.
- Do not inhale any of the fumes from the plastic.
- Before lighting the Bunsen burner, check that there are no flammable solvents nearby.

Procedure

1. Squeezing the pipette bulb, draw enough water into the pipette to form a small plug. The rest of the pipette should contain air.
2. Using a flame from a match or Bunsen burner, carefully seal the open end of the pipette completely. Allow the pipette to cool for at least 3 min before carrying on with the rest of the procedure.
3. Using scissors, cut off the tip of the bulb of the pipette.
4. Carefully attach the pipette to the ruler, using a rubber band, so that the bottom of the tube is even with the 1.0 cm mark of the ruler.
5. Fill a 400 mL beaker about two thirds full of tap water and add 3 or 4 ice cubes. Place the thermometer in the water. Then put the ruler with attached pipette into the water. Allow the ruler and pipette to sit for 5 min.



6. Copy the data table into your notebook. Your table should have at least eight rows for data.

Temp. (°C)	V (length of trapped air in mm)

7. After 5 min, measure the length (or "volume") of the trapped gas in mm. Remember that the bottom of the pipette stem is set at the 1.0 cm mark. Record these values in your data table.
8. Place the beaker on the hot plate and **slowly** heat the water in the beaker. Measure the length ("V") and temperature of the trapped gas at every 10°C to 15°C. Measure the length and temperature to a maximum of 60°C.
9. Clean the apparatus and dispose of the pipette as directed.
10. Complete the data table. Find the average of the V/T column.
11. Plot a graph of V (mm) versus T. The horizontal axis (temperature) must extend from −300°C to 100°C.
12. Draw a line of best fit. Extrapolate this line to the x-intercept.

Analysis

1. What is the independent variable in this investigation? What is the dependent variable?
2. What relationship did you notice between temperature and volume?
3. When Jacques Charles did an activity similar to this one, he obtained an x-intercept of −273°C. What is significant about this value?
4. If the value obtained by Charles was correct, what is the percentage error in your x-intercept?

Conclusion

5. In your own words, state the qualitative relationship that exists between the temperature and volume of a fixed amount of gas at constant pressure.

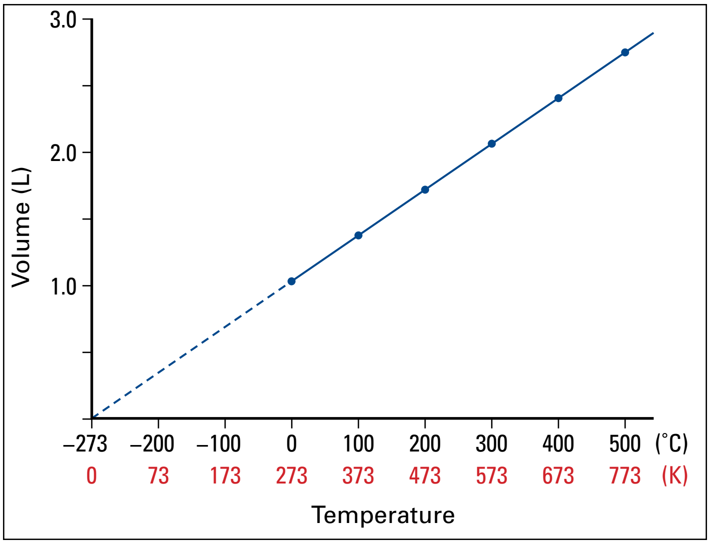


Figure 11.17 Absolute zero for an ideal gas is –273.15˚C or 0 K, the point at which all molecular motion theoretically ceases.

The Kelvin Scale and Absolute Zero

Charles found that, regardless of the gas tested, the *x*-intercept on a graph would always be –273˚C. In 1848, Lord Kelvin (1824–1907), a Scottish scientist, realised the significance of this finding. He reasoned that at –273˚C, molecular motion would cease. At this temperature, kinetic energy would be zero. The volume of a gas would, hypothetically, also be zero.

Of course, real gas molecules do have volume. Also, at low temperatures all gases will condense and change state. Still, Kelvin used this reasoning as the basis for a new temperature scale, the **Kelvin scale**. The starting point for the scale, 0 K, is called **absolute zero**. Figure 11.17 shows how absolute zero can be hypothesized, based on data from experiments. The modern accepted value for 0 K, derived with equipment more sophisticated than that available to Charles, is –273.15˚C. Each unit in the Kelvin scale is exactly the same as a unit in the Celsius scale.

There are no degree signs used in the Kelvin scale. More importantly, there are no negative values. What would happen if you tried to calculate a temperature twice as warm as –5˚C? Mathematically, the answer would be –10˚C, but this is a *colder* temperature. When mathematical manipulations are involved in studying gas behaviour, you need to convert degree Celsius temperatures into kelvins. This is done using the relationship

$$T_K = {}^{\circ}\mathrm{C} + 273.15 \textbf{ or } {}^{\circ}\mathrm{C} + 273$$

Most often, you will round off and use 273 as the conversion factor relating K and ˚C.

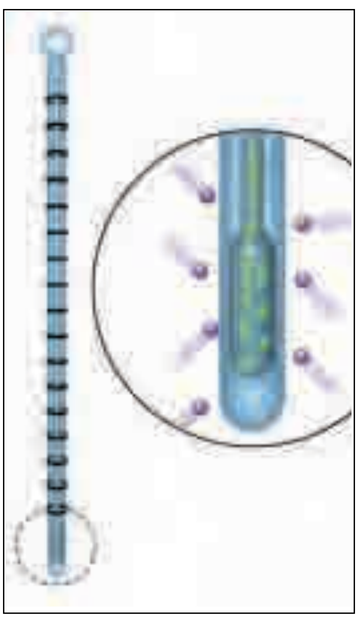


Figure 11.18 As temperature increases, particles move more rapidly, striking the outside of the thermometer with greater force and frequency. The kinetic energy of the particles is transferred to the particles inside the tube of the thermometer. The volume of the liquid inside the tube (usually mercury or coloured alcohol) expands.

Charles' Law

Although Charles discovered that the volume of a fixed amount of gas at constant pressure was proportional to its temperature, he never published this finding. In 1802, Joseph Louis Gay-Lussac (1778–1850), a French scientist, made reference to Charles' work in a published paper. The relationship between temperature and volume has since become known as **Charles' law**. Charles' law states that *the volume of a fixed mass of gas is proportional to its temperature when the pressure is kept constant*.

As you can see from Figure 11.19 on page 442, the volume of a gas increases or decreases by a fixed increment when subjected to a change in temperature. The algebraic statement of Charles' law depends on using absolute, or Kelvin, temperatures. This law is stated as $V \alpha T$, where T is measured in kelvins. (Figure 11.18 uses Charles' law to explain how a thermometer works.)

Introducing a proportionality constant (k_1), this relationship can be restated as

$$V = k_1 T \qquad \text{or} \qquad \frac{V}{T} = k_1$$

This relationship only applies if pressure is kept constant and temperature is given in kelvins. If a sample of gas is collected at initial conditions (i), this relationship can be rewritten as

$$\frac{V_i}{T_i} = k_1$$

Suppose the gas sample is subjected to a change in temperature. Under the final conditions (f), there will be a volume change such that

$$\frac{V_f}{T_f} = k_1$$

Since both initial and final conditions are equal to the proportionality constant, Charles' law can be written as

$$\frac{V_i}{T_i} = \frac{V_f}{T_f}$$

Thus, Charles' law can be restated as: *the volume of a fixed mass of gas at constant pressure is directly proportional to its Kelvin temperature.* In the next ThoughtLab, you will convert your temperature findings from Investigation 11-B into kelvins and see why you must use Kelvin temperature when performing calculations with gases.

Electronic Learning Partner

Go to the Chemistry 11 Electronic Learning Partner for a demonstration of Charles' law.

ThoughtLab Charles' Law and Kelvin Temperature

As you have learned, when making calculations involving temperature in gas samples, Kelvin temperatures must be used. You will see why for yourself in this ThoughtLab.

Materials

data table from Investigation 11-B
graph paper

Procedure

1. Make a new data table like the one below. You should include at least eight rows for data.

Temp (°C)	Temp (K)	V (mm)	$\frac{V\text{ (mm)}}{T\text{ (°C)}}$	$\frac{V\text{ (mm)}}{T\text{ (K)}}$

2. Fill in columns 1, 3, and 4 with your data from Investigation 11-B.
3. Convert temperature data in column 1 from °C to kelvins. Enter the new values in columns 2 and 5.
4. Plot a graph of V (mm) versus T (K). The horizontal axis (T) must extend from 0 K to 400 K.
5. Draw a line of best fit. Extrapolate this line to the x-intercept.

Analysis

1. What did you notice about the values of V/T (K) in the data table?
2. How do the values of V/T (K) compare to the values of V/T (°C) in the data table? Explain the significance of these two sets of data.
3. What mathematical relationship seems to exist between volume and temperature when temperature is recorded in °C? In kelvins?

Extension

4. Make a new data table like the one below. Include at least eight rows for data.

At 100 kPa		At 163 kPa		At 346 kPa	
T (K)	V (mm)	T (K)	V (mm)	T (K)	V (mm)

In columns 1 and 2, enter your data from columns 2 and 3 of the data table you made above. Carry the data from column 1 into columns 3 and 5.

5. Using the Boyle's law formula ($P_iV_i = P_fV_f$ at constant T and n), calculate the "volume" of the gas sample at 163 kPa and at 346 kPa.
6. Plot a graph of your new data table. Use a different colour for each line. Extrapolate each line to the x-intercept.
7. Of what significance are the results obtained from this graph?

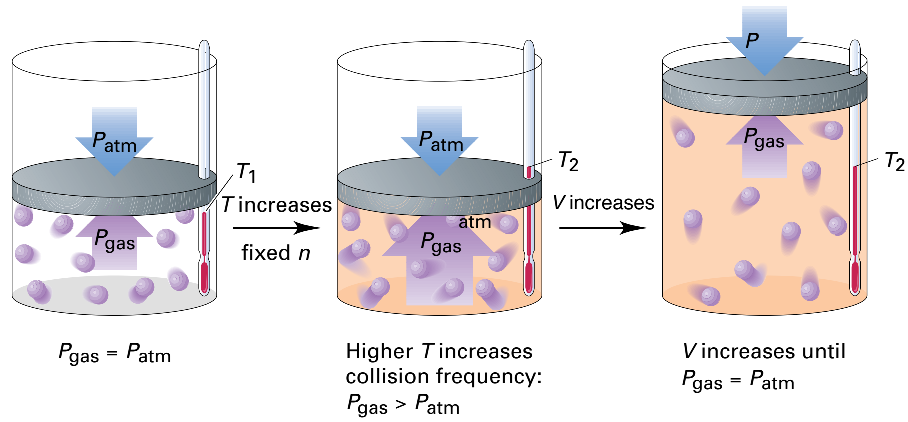


Figure 11.19 **As a gas is heated from T_i to T_f, the molecules move faster and collide with the container walls more frequently, increasing the pressure applied by the gas (P_{gas}). This added pressure increases the volume of the container until the pressure exerted by the gas is equal to the pressure exerted by the atmosphere.**

CHEM FACT

As helium gas is cooled below –268.95°C, it forms a liquid. At –270.97°C, helium still looks like a liquid, but a liquid with unusual properties. Suddenly, liquid density drops and this "liquid" gains the ability to move through very small holes that helium gas cannot pass through. It flows up the walls of its container defying gravity, and has zero viscosity. Below –270.97°C, helium becomes a superfluid, the only one discovered so far. Helium never changes to a solid.

Sample Problem

Charles' Law: Calculating Volume

Problem

Using a glass syringe, a scientist draws exactly 25.5 cm^3 of dry oxygen at 20.0°C from a metal cylinder. She needs to heat the oxygen for an experiment, so she places the syringe in an oven at 65.0°C and leaves it there for 30 min. Assuming the atmospheric pressure remains the same, what volume will the oxygen occupy?

What Is Required?

You need to find the volume of the oxygen in the syringe after it has been heated for 30 min. (V_f = ?)

What Is Given?

- You know the initial volume and temperature.
  Initial volume (V_i) = 25.5 cm^3
  Initial temperature (T_i) = 20.0°C
- You know the final temperature.
  Final temperature (T_f) = 65.0°C
- You know that the pressure does not change.

Continued ...

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FROM PAGE 442

Plan Your Strategy

Algebraic method

- Since the pressure is constant and the temperature of the gas increases, you will need to use the Charles' law formula to find the final volume of the gas sample.
- Since T is given in °C, you will need to convert it to kelvins.
- You can substitute numbers and units for the variables in the formula to solve for the unknown (V_f).

Ratio method

- Since the pressure is constant and the gas is subjected to an increase in temperature, you know that, according to Charles' law, the volume will increase.
- Since T is given in °C, you will need to convert it to kelvins.
- To find the final volume, you can multiply the initial volume by a ratio of the kelvin temperatures that is greater than one.

Act on Your Strategy

Algebraic method

$$T_i = (20.0\text{°C} + 273) = 293\text{ K} \qquad T_f = (65.0\text{°C} + 273) = 338\text{ K}$$

$$\frac{V_i}{T_i} = \frac{V_f}{T_f}$$

$$\frac{25.5\text{ cm}^3}{293\text{ K}} = \frac{V_f}{338\text{ K}}$$

To isolate V_f, you need to multiply both sides of the equation by 338 K.

$$\frac{25.5\text{ cm}^3}{293\text{ K}} \times 338\text{ K} = \frac{V_f}{\cancel{338\text{ K}}} \times \cancel{338\text{ K}}$$

$$\frac{(25.5\text{ cm}^3)(338\ \cancel{\text{K}})}{(293\ \cancel{\text{K}})} = V_f$$

$$V_f = 29.42\text{ cm}^3$$

Ratio method

$$T_i = (20.0\text{°C} + 273) = 293\text{ K} \qquad T_f = (65.0\text{°C} + 273) = 338\text{ K}$$

$$V_f = V_i \times \text{ temperature ratio}$$
$$= 25.5\text{°C} \times \frac{338\ \cancel{\text{K}}}{293\ \cancel{\text{K}}}$$
$$= 29.42\text{ cm}^3$$

Continued ...

PROBLEM TIP

In these Sample Problems, you will see two different methods of solving the problem: the algebraic method and the ratio method. Choose the method you prefer to solve this type of problem.

Continued ...
FROM PAGE 443

Since the least number of significant digits in the question is three, the final volume will be reported to three significant digits.

$$V_f = 29.4\ \text{cm}^3$$

Check Your Solution

- The units for the answer are in cubic centimetres.
- When the units cancel out, cm^3 remains.
- The volume of the oxygen gas increased due to an increase in temperature.

Sample Problem

Charles' Law: Calculating Temperature

Problem

A balloon is filled with 2.50 L of dry helium at 23.5°C. The balloon is placed in a freezer overnight. The next morning, the balloon is removed and the volume is found to be 2.15 L. What was the temperature (in °C) inside the freezer if the pressure remained constant?

What Is Required?

You need to find the temperature of the freezer in °C. (T_f = ?)

What Is Given?

- You know the initial volume and temperature.
  Initial volume (V_i) = 2.50 L
  Initial temperature (T_i) = 23.5°C
- You know the final volume.
  Final volume (V_f) = 2.15 L
- You know that the pressure does not change.

Plan Your Strategy

Algebraic method

- Since pressure is constant and the volume and temperature change, you will need to use the Charles' law formula to find the final temperature of the gas sample.
- Since T is given in °C, you need to convert it to kelvins.
- You can substitute numbers and units for the variables in the formula to solve for the unknown (T_f).

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FROM PAGE 444

Ratio method

- Since pressure remains constant and the volume of the balloon decreases, you know that according to Charles' law, the temperature of the gas must also decrease.
- Since T is given in °C, you need to convert it to kelvins.
- To find the final temperature inside the freezer, you can multiply the initial temperature by a volume ratio that is less than one.

Act on Your Strategy

Algebraic method

$$T_i = (23.5^\circ\text{C} + 273)$$
$$= 297\text{ K}$$

$$\frac{V_i}{T_i} = \frac{V_f}{T_f}$$

$$\frac{2.50\text{ L}}{297\text{ K}} = \frac{2.15\text{ L}}{T_f}$$

To simplify the equation and make it easier to solve, you can first cross-multiply the above equation.

$$(2.50\text{ L})(T_f) = (296.5\text{ K})(2.15\text{ L})$$

To isolate T_f, you need to divide both sides of the equation by 2.50 L.

$$\frac{(\cancel{2.50\text{ L}})(T_f)}{(\cancel{2.50\text{ L}})} = \frac{(297\text{ K})(2.15\text{ L})}{(2.50\text{ L})}$$

$$T_f = \frac{(297\text{ K})(2.15\ \cancel{\text{L}})}{(2.50\ \cancel{\text{L}})}$$
$$= 255.42\text{ K}$$

Since the question asks for the temperature in °C, you need to convert kelvins to °C. To do this, subtract 273 from the answer.

$$T_f = (255.42\text{ K} - 273)$$
$$= -17.6^\circ\text{C}$$

Ratio method

$$T_i = (23.5^\circ\text{C} + 273)$$
$$= 297\text{ K}$$

$$T_f = 297\text{ K} \times \text{volume ratio}$$
$$= 297\text{ K} \times \frac{2.15\ \cancel{\text{L}}}{2.50\ \cancel{\text{L}}}$$
$$= 255.42\text{ K}$$

Since the question asks for the temperature in °C, you need to convert kelvins to °C. To do this, subtract 273 from the answer.

$$T_f = (255.42\text{ K} - 273)$$
$$= -17.6^\circ\text{C}$$

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FROM PAGE 445

mind STRETCH

Using the formula $y = mx + b$, see if you can derive the Charles' law formula from the graph that you produced in the ThoughtLab on page 441. For help in deriving equations from graphs, see Appendix E.

Since the least number of significant digits in the question is three significant digits, the final temperature will be reported to the same significant digits.

$T_f = -17.6^\circ\text{C}$

Check Your Solution

- The unit for the answer is in kelvins.
- When the units cancel out, kelvins remain.
- Kelvins have been converted to °C.
- The temperature of the balloon has decreased, which is reflected in its decrease in volume.

Practice Problems

5. Convert the following temperatures to the Kelvin scale.

(a) 25°C

(b) 37°C

(c) 150°C

6. Convert the following temperatures to degrees Celcius.

(a) 373 K

(b) 98 K

(c) 425 K

7. Give an example of something that might be at each temperature in question 5.

8. A sample of nitrogen gas surrounding a circuit board occupies a volume of 300 mL at 17°C and 100 kPa. What volume will the nitrogen occupy at 100.0°C if the pressure remains constant?

9. A 2.5 L balloon is completely filled with helium indoors at a temperature of 24.2°C. The balloon is taken out on a cold winter day (–17.5°C). What will the volume of the balloon become, assuming a constant pressure?

10. 10.0 L of neon at 20.0°C is expanded to a volume of 30.0 L. If the pressure remains constant, what must the final temperature be (in °C)?

11. A 14.5 cm^3 sample of oxygen gas at 24.3°C is drawn into a syringe with a maximum volume of 60 cm^3. What is the maximum change in temperature that the oxygen can be subjected to before the plunger pops out of the syringe?

12. Methane gas can be condensed by cooling and increasing the pressure. A 600 L sample of methane gas at 25°C and 100 kPa is cooled to –20°C. In a second step, the gas is compressed until the pressure is quadrupled. What will the final volume be? (**Hint:** Use both Boyle's law and Charles' law to answer this question.)

PROBLEM TIP

In question 8, the smallest number of significant digits is two (17°C). However, before doing your calculations, you must convert this value to kelvins (290 K). This value now has *three* significant digits. Round off your final answer to *three* significant digits.

Gay-Lussac's Law

Aside from balloons and syringes, most containers that are used to store gases have a fixed volume. You know that temperature is a measure of the average kinetic energy of the molecules making up a substance. If the temperature of a gas increases, but the volume of its container cannot increase, what happens to the pressure of the gas inside?

Extending the work of Charles, Joseph Louis Gay-Lussac discovered the relationship between temperature and pressure acting on a fixed volume of a gas. (Remember that for gases, V_{gas} = volume of container holding the gas.) As you will learn later in this section, this relationship is very important for the safe handling of gases under pressure in steel tanks or aerosol cans. **Gay-Lussac's law** states that *the pressure of a fixed amount of gas, at constant volume, is directly proportional to its Kelvin temperature*. (See Figure 11.20.)

$P \propto T$

(if T is given in kelvins and volume and amount of gas is constant)

Introducing a new proportionality constant (k_2), this relationship can be restated as

$$P = k_2 T \qquad \text{or} \qquad \frac{P}{T} = k_2$$

If we assign P_i and T_i as the initial conditions, and P_f and T_f for the final conditions, the above relationship can be rewritten as

$$\frac{P_i}{T_i} = \frac{P_f}{T_f}$$

As you will notice, this mathematical relationship is very similar to that of Charles' law.

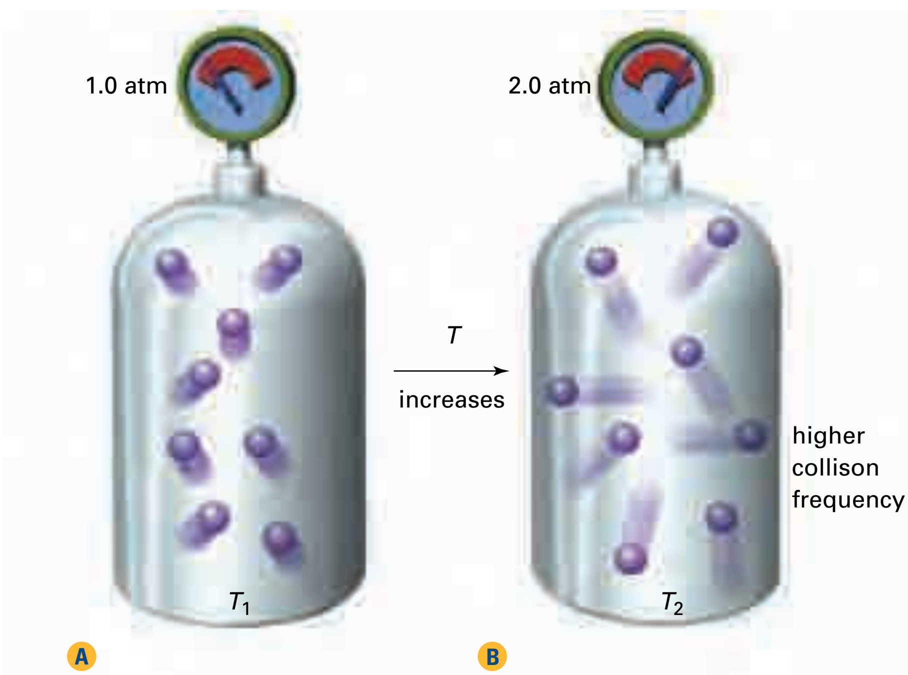

Figure 11.20 **When temperature increases from T_i to T_f in a rigid container with a constant volume, the average speed of the gas molecules increases. Since the molecules move faster, they collide with each other and with the walls of the container more forcefully and more frequently. The gas pressure increases.**

Technology LINK

Underinflated vehicle tires contribute to unsafe road handling and to lower fuel economy. Based on what you know about Gay-Lussac's law, why should you measure the pressure in vehicle tires before driving the vehicle for a long distance?

Web LINK

www.school.mcgrawhill.ca/resources/

Refrigerators and air conditioners function because of the relationship between pressure and temperature. Research the Joule-Thomson Effect. Go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next.

Sample Problem

Gay-Lussac's Law: Calculating Pressure

Problem

A cylinder of chlorine gas (Cl_2) is stored in a concrete-lined room for safety. The cylinder is designed to withstand 50 atm of pressure. The pressure gauge reads 35.0 atm at 23.2°C. An accidental fire in the room next door causes the temperature in the storage room to increase to 87.5°C. What will the pressure gauge read at this temperature?

What Is Required?

You need to find the pressure of the oxygen once the temperature has been increased. (T_f = ?)

What Is Given?

- You know the initial pressure and temperature.
  Initial pressure (P_i) = 35.0 atm
  Initial temperature (T_i) = 23.2°C
- You know the final temperature.
  Final temperature (T_f) = 87.5°C
- You know that the volume of the rigid metal tank will not change appreciably.

PROBLEM TIP

In these Sample Problems, you will see two different methods of solving the problem: the algebraic method and the ratio method. Choose the method you prefer to solve this type of problem.

Plan Your Strategy

Algebraic method

- Since the volume of the cylinder is essentially constant, and the temperature increases, you will need to use the Gay-Lussac's law formula to find the final pressure of the gas sample.
- Since T is given in °C, you need to convert it to kelvins.
- You can substitute numbers and units for the variables in the formula to solve for the unknown. (P_f = ?)

Ratio method

- Since the volume of the cylinder is essentially constant, and the temperature increases, you know that according to Gay-Lussac's law, the pressure exerted by the gas increases as well.
- Since T is given in °C, you need to convert it to kelvins.
- To find the final pressure of the gas, you can multiply the initial pressure by a temperature ratio that is greater than one.

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Act on Your Strategy

Algebraic method

$T_i = (23.2°C + 273) = 296 \text{ K}$ $T_f = (87.5°C + 273) = 360 \text{ K}$

$$\frac{P_i}{T_i} = \frac{P_f}{T_f}$$

$$\frac{35.0 \text{ atm}}{296 \text{ K}} = \frac{P_f}{361 \text{ K}}$$

To simplify the equation and make it easier to solve, you can first cross-multiply the above equation.

$(35.0 \text{ atm})(361 \text{ K}) = (P_f)(296 \text{ K})$

To isolate P_f, you need to divide both sides of the equation by 296 K.

$$\frac{(35.0 \text{ atm})(361 \text{ K})}{(296 \text{ K})} = \frac{(P_f)\cancel{(296 \text{ K})}}{\cancel{(296 \text{ K})}}$$

$$\frac{(35.0 \text{ atm})(361 \cancel{\text{K}})}{(296 \cancel{\text{K}})} = P_f$$

$$P_f = 42.69 \text{ atm}$$

Ratio method

$T_i = (23.2°C + 273) = 296 \text{ K}$ $T_f = (87.5°C + 273) = 361 \text{ K}$

$$\begin{aligned} P_f &= 35.0 \text{ atm} \times \text{temperature ratio} \\ &= 35.0 \text{ atm} \times \frac{361 \cancel{\text{K}}}{296 \cancel{\text{K}}} \\ &= 42.69 \text{ atm} \end{aligned}$$

Since the least number of significant digits in the question is three, the final pressure will be rounded off to the same number of significant digits.

$$P_f = 42.7 \text{ atm}$$

Check Your Solution

- The unit for the answer is in atmospheres.
- When the units cancel out, atm remains.
- Kelvins have been converted to °C.
- The pressure inside the cylinder has increased, as would be expected when the temperature increases.

Practice Problems

13. An unknown gas is collected in a 250.0 mL flask and sealed. Using electronic devices, it is found that the gas inside the flask exerts a pressure of 135.5 kPa at 15°C. What pressure will the gas exert if the temperature (in Kelvins) is doubled?

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14. At 18°C, a sample of helium gas stored in a metal cylinder exerts a pressure of 17.5 atm. What will the pressure become if the tank is placed in a closed room where the temperature increases to 40°C?

15. A gaseous refrigerant, enclosed in copper tubes, surrounds the freezer in a small refrigerator. The gas is found to exert a pressure of 110 kPa at 45°C. The refrigerant is allowed to expand through a nozzle into an expansion chamber such that the exerted pressure decreases to 89 kPa. What is the temperature inside the freezer?

16. Before leaving on a trip to Florida, you measure the pressure inside the tires of your car at a gas station. At −7.5°C the tire pressure is found to be 206.5 kPa. When you arrive in Florida, you stop for dinner. Before leaving, you once again measure the tire pressure at a gas station beside the restaurant. Most pressure gauges in the United States are calibrated in psi. You find the tire pressure to be 34.3 psi. What is the approximate temperature in Florida? (**Hint:** See the MathLink on page 428 to find out how to convert psi to kPa.)

Compressed Gases and Safety Concerns

The Gay-Lussac's law Sample Problem in this section indicates how carefully gases under pressure must be handled. Chlorine gas can cause serious respiratory problems and irritate the skin and mucous membranes. In extreme cases, death from suffocation could result from exposure to this gas. Yet chlorine is an important industrial product. Compounds of chlorine are used in bleaches, oxidizing agents, and solvents, and as intermediates in the manufacture of other substances.

Compressed gases are commonly stored in thick-walled metal cylinders designed especially for this purpose. All cylinders must comply with Canadian Transport Commission (CTC) regulations. Containers must be permanently marked with a serial number and specifications for the volume of the cylinder and the maximum pressure it can withstand. Containers must be tested every five to ten years, with the date of the test stamped on the cylinder.

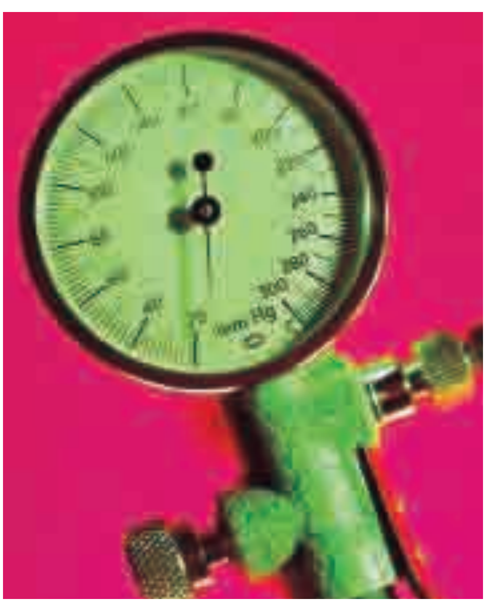

Figure 11.21 **Gas cylinders have pressure valves to help the user regulate the amount of gas escaping when the cylinder is opened.**

Figure 11.22 on the next page shows a typical compressed gas cylinder. You can see that it is built to withstand high pressures. There are other safety precautions as well, however.

Most cylinders used to store gases have safety devices regulating the internal gas pressure. The most common of these is a **pressure relief valve**. If the pressure inside the cylinder increases to a dangerous level, a spring allows the valve to open and release excess gas until the internal pressure returns to a safe level. Some pressure relief valves will close once excess gas is released. These valves are relatively expensive compared to non-reclosing valves. Non-reclosing valves are found on common household products such as aerosol hairsprays.

Cylinders used to store gases such as acetylene (C_2H_2), that could cause an explosive chemical reaction at high temperatures, are fitted with **fusible plugs**. These plugs are designed to melt and allow gas to escape at temperatures lower than those at which hazardous reactions can start. Fusible plugs for acetylene cylinders are made of a metal alloy that melts at 100°C.

Not all compressed gas cylinders are fitted with pressure-relief valves or fusible plugs. Cylinders containing toxic gases such as chlorine or phosgene ($COCl_2$) are one example. These gases could cause serious harm to health if released into the air in sufficient quantities. Therefore, these gases, like all compressed gases, must be handled with great care. They must be stored in a well-ventilated, dry area. The surrounding storage area must be fire-resistant, and proper fire-fighting equipment must be immediately available. Gas cylinders should never be stored near electrical circuits that might spark or near other ignition sources such as open flames. Material Data Safety Sheets must be available, and all cylinders must be clearly labelled with WHMIS warning signs.

Figure 11.22 **Pressure relief valves prevent a compressed gas cylinder from exploding if the temperature, and thus the pressure increases.**

Section Wrap-up

In this section, you learned about the relationship between volume and temperature (Charles' law). You also learned about the relationship between pressure and temperature (Gay-Lussac's law). In the next section, you will see how these relationships can be combined with Boyle's law to produce one equation that works in all three situations.

Section Review

1. K/U What safety concerns and precautions should be taken with compressed gases? Use what you know about the movement of particles to explain these precautions.

2. **(a)** I A gas at 107 kPa and 300 K is cooled to 146 K at the same volume. What is the new pressure?

   (b) I 17 L of gas at 300 K are cooled to 146 K at the same pressure. What is the new volume?

3. MC Describe the relationship between your answers in parts (a) and (b) of question 2.

4. C Explain in terms of molecular motion why, when the temperature is increased:

   (a) a gas increases in volume

   (b) the pressure of a gas increases

5. K/U A balloon at a party drifts above a hot stove, and explodes. Why did this happen?

11.4 Combined Gas Law Calculations

Section Preview/ Specific Expectations

In this section, you will

- **solve** problems involving the combined gas law and Dalton's law of partial pressures
- **identify** the components of Earth's atmosphere
- **communicate** your understanding of the following terms: *standard temperature, standard temperature and pressure, standard ambient temperature and pressure, combined gas law, Dalton's law of partial pressures*

You may have heard a common joke about Canadian weather: "If you don't like it, wait an hour and it will change." While this is an exaggeration, atmospheric pressure and temperature rarely remain constant for any extended period of time. Since the volume of gases changes when pressure and temperature change, standards have been designed to allow a comparison of different gas volumes.

The average pressure of the atmosphere at sea level is taken as standard pressure (760 mm Hg = 760 torr = 1 atm = 101.3 kPa). The freezing point of water (0˚C or 273 K) is defined as **standard temperature**. Together, these conditions are referred to as **standard temperature and pressure** (**STP**). (See Figure 11.23.) The normal conditions under which we live are referred to as **standard ambient temperature and pressure**. These conditions are known as **SATP**, defined as 25˚C and 100 kPa.

How could we find what the volume of a gas, measured under different conditions, would be when changed to STP or SATP?

Figure 11.23 This photograph illustrates typical STP conditions: 0˚C at sea level.

Boyle established that pressure and volume are inversely proportional:

$$P_iV_i = P_fV_f$$

Charles found that volume and temperature are directly proportional:

$$\frac{V_i}{T_i} = \frac{V_f}{T_f}$$

Gay-Lussac discovered that pressure has the same relationship to temperature as volume does:

$$\frac{P_i}{T_i} = \frac{P_f}{T_f}$$

Do you notice a pattern here?

Pressure and volume are directly related to temperature, and inversely related to each other. Write this as one law, and it is possible to calculate situations in which three variables change at the same time. The mathematics work out just as consistently as in the two-variable equations.

CHECKPOINT

Under what conditions might you use standard ambient temperature and pressure as a reference rather than standard temperature and pressure ?

The Combined Gas Law

Boyle's law can be used to solve for changes in volume when pressure changes. The gas must be in a closed system and the temperature must remain constant. You can use Charles' law to solve for changes in volume with temperature changes. This law works only in a closed system in which pressure remains constant. Gay-Lussac's law can solve problems in which the amount and volume of gas remain constant while the temperature and pressure change.

As you learned, Boyle's law ($V \propto 1/P$) is expressed mathematically as $PV = k$. Charles' law ($V \propto T$) is expressed mathematically as $V/T = k_1$ when temperature is recorded in kelvins. Combining these two expressions gives:

$$V \propto \frac{1}{P} \times T \qquad \text{or} \qquad V \propto \frac{T}{P}$$

Introducing a new proportionality constant (k_3), we can write

$$V = \frac{T}{P} \times k_3 \qquad \text{or} \qquad \frac{PV}{T} = k_3$$

This mathematical relationship is the **combined gas law**.

If a sample of a gas is trapped at a measured set of initial conditions, the combined gas law can be rewritten as

$$\frac{P_iV_i}{T_i} = k_3$$

As this gas is then subjected to changes in pressure and temperature, the final condition of the gas can be described mathematically as

$$\frac{P_fV_f}{T_f} = k_3$$

Since both expressions are equal to the same proportionality constant (k_3), the combined gas law can be written as

$$\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f}$$

Sample Problem

Finding Volume: The Combined Gas Law

Problem

Sandra is having a birthday party on a mild winter's day. The weather changes and a higher-pressure (103.0 kPa) cold front (–25˚C) rushes into town. The original air temperature was –2˚C and the pressure was 100.8 kPa. What will happen to the volume of the 4.2 L balloons that were tied to the front of the house?

What Is Required?

You need to find the volume of the balloons under the new conditions of temperature and pressure. (V_f = ?)

Continued ...

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What Is Given?

- You know the initial pressure, volume, and temperature.
  Initial pressure $(P_i) = 100.8$ kPa
  Initial volume $(V_i) = 4.2$ L
  Initial temperature $(T_i) = -2°C$
- You know the final pressure and temperature.
  Final pressure $(P_f) = 103.0$ kPa
  Final temperature $(T_f) = -25°C$

Plan Your Strategy

Algebraic method

- Since pressure and temperature both change, use the combined gas law to find the final volume of the balloon.
- Since T is given in °C, you need to convert it to kelvins.
- You can rearrange the combined gas law and substitute numbers and units for the variables in the formula to solve for V_f.

Ratio method

- Since pressure and temperature change, you know that the volume of the balloon will also change.
- Since T is given in °C, you need to convert it to kelvins.
- To find the new volume based on the increase in pressure, you need to multiply the initial volume by a pressure ratio that is less than one.
- To find the new volume based on the decrease in temperature, you need to multiply the initial volume by a temperature ratio that is less than one.
- To find the new volume based on both pressure and temperature changes, you can multiply the initial volume by the pressure and temperature ratios.

PROBLEM TIP

In these Sample Problems, you will see two different methods of solving the problem: the algebraic method and the ratio method. Choose the method you prefer to solve this type of problem.

Act on Your Strategy

Algebraic method

$T_i = (-2°C + 273) = 271$ K $\qquad T_f = (-25°C + 273) = 248$ K

$$\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f}$$

Solve the combined gas law, in this case for V_f.

Continued ...

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To isolate V_f, move T_f and P_f. Multiply T_f up to the numerator on the left side, and divide P_f down to the denominator. This leaves

$$\frac{P_i V_i T_f}{T_i P_f} = V_f$$

$$\frac{(100.8\ \cancel{\text{kPa}})(4.2\ \text{L})(248\ \cancel{\text{K}})}{(271\ \cancel{\text{K}})(103.0\ \cancel{\text{kPa}})} = V_f$$

$$V_f = 3.76\ \text{L}$$

$$\text{So} \quad V_f = 3.8\ \text{L}$$

Ratio method

$T_i = (-2°\text{C} + 273) = 271\ \text{K}$ $\qquad$ $T_f = (-25°\text{C} + 273) = 248\ \text{K}$

$V_f = V_i \times$ pressure ratio $\times$ temperature ratio

$$= 4.2\ \text{L} \times \frac{100.8\ \cancel{\text{kPa}}}{103.0\ \cancel{\text{kPa}}} \times \frac{248\ \cancel{\text{K}}}{271\ \cancel{\text{K}}}$$

$$= 3.76\ \text{L} \cong 3.8\ \text{L}$$

Remember to change the answer to correct significant digits. The least number of digits in the question was two, so the answer must have only two significant digits. $V_f = 3.8$ L.

Check Your Solution

- The unit for the answer is in litres.
- When the units cancel out, L remains.
- The volume of the balloon has decreased, as would be expected when pressure increases and temperature decreases.

Sample Problem

Combined Gas Law

Problem

An automated instrument has been developed to help drug-research chemists determine the amount of nitrogen in a compound. Any compound containing carbon, nitrogen, and hydrogen is reacted with copper(II) oxide to produce CO_2, H_2O, and N_2 gases. The gases are collected separately and analyzed.

In an analysis of 39.8 mg of caffeine using this instrument, 10.1 mL of N_2 gas is produced at 23°C and 746 torr. What must the new temperature of nitrogen be, in °C, if the volume is increased to 12.0 mL, and the pressure is increased to 780 torr?

What Is Required?

You need to find the temperature of the nitrogen under the new conditions of volume and pressure. ($T_f = ?$)

Continued ...

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FROM PAGE 455

What Is Given?

- You know the initial pressure, volume, and temperature.
  Initial pressure (P_i) = 746 torr
  Initial volume (V_i) = 10.1 mL
  Initial temperature (T_i) = 23°C
- You know the final pressure and volume.
  Final pressure (P_f) = 780 torr
  Final volume (V_f) = 12.0 mL

Plan Your Strategy

Algebraic method

- Since pressure and volume both change, you will need to use the combined gas law formula to find the final temperature of the nitrogen.
- Since T is given in °C, you need to convert it to kelvins.
- You can rearrange the combined gas law and substitute numbers and units for the variables in the formula to solve for T_f.

Ratio method

- Since pressure and volume change, you know that the temperature of the nitrogen will also change.
- Since T is given in °C, you need to convert it to kelvins.
- To find the new temperature based on the increase in pressure, you need to multiply the initial volume by a pressure ratio that is greater than one.
- To find the new temperature based on an increase in volume, you need to multiply the initial temperature by a volume ratio that is less than one.
- To find the new temperature based on both pressure and volume changes, you can multiply the initial temperature by the pressure and volume ratios.

Act on Your Strategy

Algebraic method

$T_i = (23°C + 273) = 296 \text{ K}$

$$\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f}$$

Divide both sides by P_fV_f to isolate T_f.

$$\frac{P_iV_i}{T_iP_fV_f} = \frac{1}{T_f}$$

Now flip both sides of the equation over:

$$\frac{T_iP_fV_f}{P_iV_i} = \frac{T_f}{1}$$

Continued ...

Continued ...
FROM PAGE 456

Alternatively, you could have divided down P_iV_i and multiplied both T_f and T_i up. This is an equally correct procedure and the result would be the same. Put in the numbers:

$$\frac{(296\text{ K})(780\ \cancel{\text{torr}})(12.0\ \cancel{\text{ml}})}{(746\ \cancel{\text{torr}})(10.1\ \cancel{\text{ml}})} = T_f$$

$$T_f = 368\text{ K}$$

Since the question asks for the temperature in °C, you need to convert kelvins to °C. To do this, subtract 273 from the answer.

$$T_f = (368 - 273)$$
$$= 95°\text{C} \qquad \text{(Note the correct number of significant digits.)}$$

Ratio method

$$T_i = (23°\text{C} + 273) = 296\text{ K}$$

$$T_f = T_i \times \text{pressure ratio} \times \text{temperature ratio}$$
$$= 296\text{ K} \times \frac{780\text{ torr}}{746\text{ torr}} \times \frac{12.0\text{ mL}}{10.1\text{ mL}}$$
$$= 368\text{ K}$$

Since the question asks for the temperature in °C, you need to convert kelvins to °C. To do this, subtract 273 from the answer.

$$T_f = (368 - 273)$$
$$= 95°\text{C} \qquad \text{(Note the correct number of significant digits.)}$$

Check Your Solution

- The unit for the answer is in degrees Celcius.
- The temperature of the nitrogen has increased, as would be expected when pressure and volume increase.

Practice Problems

17. A sample of gas has a volume of 150 mL at 260 K and 92.3 kPa. What will the new volume be at 376 K and 123 kPa?

18. A cylinder at 48 atm pressure and 290 K releases 35 mL of carbon dioxide gas into a 4.0 L container at 297 K. What is the pressure inside the container?

19. In a large syringe, 48 mL of ammonia gas at STP is compressed to 24 mL and 110 kPa. What must the new temperature of the gas be?

20. A 100 W light bulb has a volume of 180.0 cm^3 at STP. The light bulb is turned on and the heated glass expands slightly, changing the volume of the bulb to 181.5 cm^3 with an internal pressure of 214.5 kPa. What is the temperature of the light bulb (in °C)?

21. Sulfur hexafluoride, $SF_{6(g)}$, is used as a chemical insulator. A 5.0 L sample of this gas is collected at 205.0°C and 350 kPa. What pressure must be applied to this gas sample to reduce its volume to 1.7 L at 25°C?

Gases and Natural Phenomena

Gases under pressure are responsible for natural phenomena such as volcanoes and geysers. On March 20, 1980, residents in the northern part of Washington state heard a rumble from the mountains. They were told it was only a minor earthquake. By March 27, seismologists were sure that something more was involved. Deep inside one of the mountains, Earth's crust was moving. The lower portion of the crust was melting into hot liquid called *magma*. The magma started rising up through cracks in the crust. Trapped water quickly turned into superheated steam. Trapped gases added to the increasing pressure inside the mountain. Then, on May 18, 1980, the top half of Mount St. Helens blew away in a gigantic explosion. It released all of the built-up pressure, along with many tonnes of rock and ash, twenty kilometres into the atmosphere. It also released steam, and gases such as carbon dioxide, nitrogen, and sulfur dioxide. (See Figure 11.24.)

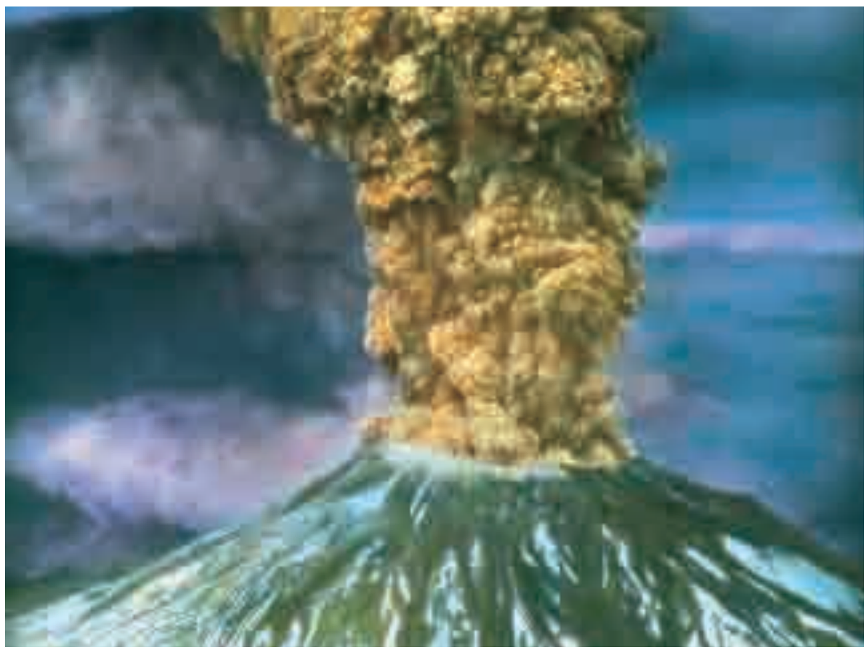


Figure 11.24 **Before a volcano erupts, there is a tremendous build-up of fluid and gas pressures inside the volcano due to magma, steam, and gases.**

Other gases released in volcanic explosions include the oxides, sulfides, and chlorides of carbon, sulfur, hydrogen, chlorine, and fluorine. These include CO_2, CO, SO_2, SO_3, H_2, H_2S, Cl_2, HCl, and F_2. It is estimated that the eruption of Mt. Pinatubo in the Philippines on June 15, 1991 released about 1.8×10^{10} kg of gases and ash into the atmosphere.

After a volcanic blast, or in an area of volcanic activity, geysers may form. Cool water from Earth's surface trickles down between rocks. It drains deep into Earth's crust, into regions where hot magma is still present. As the water is heated, it forms steam. The steam rises back up though fissures and cracks, meeting more water and heating it up as well. More and more boiling water accumulates, until suddenly, through a narrow opening, the pressurized hot water is violently ejected. As the water cools in the air, it falls back and the process starts again. Yellowstone National Park in the United States has one of the most famous geysers in the world, *Old Faithful*. It erupts approximately once every sixty-five minutes. (Figure 11.25 shows Old Faithful erupting.)

Figure 11.25 **Geysers form in areas of early volcanic activity or after a volcanic blast. They should be approached with caution since the hot water that erupts can cause severe burns.**

In Chapter 12, you will learn more about the gas chemistry of our atmosphere.

Dalton's Law of Partial Pressures

What if there are two gases in one container, or as in the case of the atmosphere, a mixture of many gases? How does this affect the pressure?

The English scientist John Dalton did a very thorough analysis of the atmosphere. He concluded that it comprised about 79% nitrogen and 21% oxygen. (See Table 11.2 to find out how close he really was.) Dalton noticed that the water vapour, however, seemed quite variable, so he did further experiments. He obtained some very dry air, and measured the pressure in the container. He then introduced some water vapour. The pressure increased! Dalton repeated and adjusted his experiment time after time, always with the same results. He concluded that *the total pressure of a mixture of gases is the sum of the pressures of each of the individual gases.* This is called **Dalton's Law of partial pressures**. (See Figure 11.27 below.)

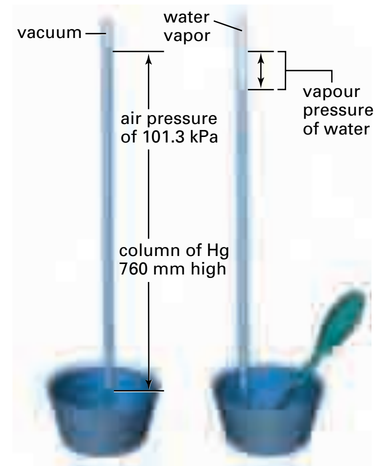


Figure 11.26 A method of determining water vapour pressure. A drop of water vapourizes in the vacuum and the water vapour exerts a pressure on the mercury in the tube.

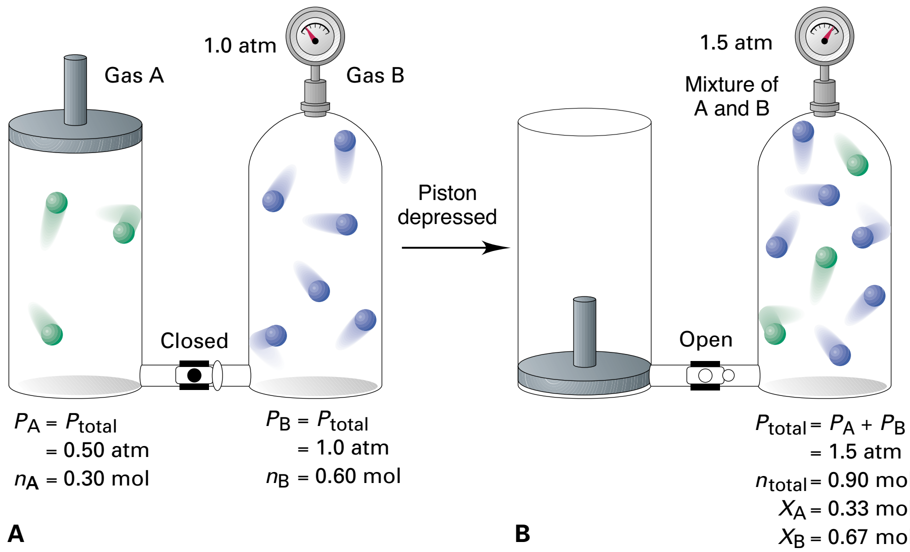


Figure 11.27 Gas A is stored at 0.5 atm and gas B at 1.0 atm. When gas A is added to gas B, the total pressure exerted is now 1.5 atm while the volume of the flask remains the same.

Table 11.2 Components of the Dry Atmosphere

Components	Percentage
Nitrogen (N_2)	78.08
Oxygen (O_2)	20.95
Argon (Ar)	0.93
Carbon dioxide (CO_2)	0.03
Neon (Ne)	0.002
Other gases	0.008

The action of winds mixes the atmosphere so that the composition of the dry atmosphere is fairly constant over the entire Earth. Water vapour, though an important component of the atmosphere, is not listed in the table. The quantity of water vapour in the air, or *humidity*, is variable. In desert climates, the quantity of water vapour will be very small (low humidity). In tropical areas, or near large bodies of water, the quantity of water vapour in the air can be quite substantial (high humidity).

Dalton extended this idea to enhance what he had discovered about the composition of the atmosphere. If the atmosphere is 79% nitrogen, and the atmospheric pressure on a certain day is, for example, 101.3 kPa, then he concluded that the nitrogen itself must be contributing

$$\frac{79}{100} \times 101.3 \text{ kPa} = 80 \text{ kPa}$$

The other gases in the atmosphere contribute pressures corresponding to their percentage of the total composition of air.

The generalized form of Dalton's law of partial pressures is

$$P_{total} = P_1 + P_2 + P_3 + \ldots + P_n$$

This law can be applied to any mixture of gases. Study the following Sample Problem and try the Practice Problems to verify your understanding of how Dalton's law works.

Sample Problem

Applying Dalton's Law of Partial Pressures

Problem

What is the pressure contribution of CO_2 to the atmospheric pressure on a very dry day when the barometer read 0.98 atm? Convert your answer into three different units.

What Is Required?

You must find the contribution of CO_2 to the atmospheric pressure, and then convert the units.

What Is Given?

- According to Table 11.2, CO_2 contributes 0.03% of the atmospheric pressure. The total atmospheric pressure is 0.98 atm. The conversion factors of pressure are
  1 atm = 760 torr = 760 mm Hg = 101.3 kPa

Plan Your Strategy

- You need to multiply the total atmospheric pressure times CO_2's contribution to the total.
- Then you can multiply the answer in atm by the conversion factors to cancel out atm and obtain other units.

Act on Your Strategy

$$\frac{0.03}{100} \times 0.98 \text{ atm} = 2.9 \times 10^{-4} \text{ atm}$$

$$2.9 \times 10^{-4} \cancel{\text{atm}} \times \frac{101.3 \text{ kPa}}{1 \cancel{\text{atm}}} = 0.030 \text{ kPa}$$

$$2.9 \times 10^{-4} \cancel{\text{atm}} \times \frac{760 \text{ torr}}{1 \cancel{\text{atm}}} = 0.22 \text{ torr} = 0.22 \text{ mm Hg}$$

Check Your Solution

- The answers given are all very small. This is expected, as CO_2 only contributes a small percentage of gas molecules to the atmosphere.

Practice Problems

22. To speed up a reaction in a vessel pressurized at 98.0 kPa, a chemist added 202.65 kPa of hydrogen gas. What was the resulting pressure?

23. A gas mixture contains 12% Ne, 23% He, and 65% Rn. If the total pressure is 116 kPa, what is the partial pressure of each gas?

24. The partial pressure of argon gas, making up 40% of a mixture, is 325 torr. What is the total pressure of the mixture in kPa?

25. A mixture of nitrogen and carbon dioxide gas is at a pressure of 1.00 atm and a temperature of 278 K. If 30% of the mixture is nitrogen, what is the partial pressure of the carbon dioxide?

Molar Mass and Gas Behaviour

In the following ThoughtLab, you will examine one more factor affecting gas behaviour:

ThoughtLab Boiling Points of Gases

What effect does molar mass have on the behaviour of a gas? In fact, molar mass helps to determine whether a compound or element is gaseous or not. You will examine the connection between boiling point and molar mass here.

Table 11.3 The Boiling Points of Various Gases

Gas	Boiling Point (°C)
He	−269
Ne	−246
N_2	−196
Ar	−186
O_2	−183
Kr	−153
Xe	−108

Materials

graph paper; periodic table

Procedure

1. Using a periodic table, record the molar mass of each of the gases listed in the table.
2. Plot a graph of boiling points of gases (y-axis) versus their molar masses (x-axis).

Analysis

1. From the graph obtained, what relationship exists between the molar mass of a gas and its boiling point?
2. Using the graph, what would be the boiling points of gases such as hydrogen (H_2), fluorine (F_2) and radon (Rn)?
3. Predict whether a substance such as bromine (Br_2) would be a gas or liquid at room temperature. Use a reference book to check your answer.

Section Wrap-up

In this section, you learned how to use the combined gas law for gas calculations. You also learned about natural phenomena that are related to gases. Finally, you learned about Dalton's law of partial pressures. In Chapter 12, you will learn more about gas laws. First, however, you will take a closer look at some technological applications of the gas laws.

Section Review

1. K/U In a large syringe, 48 mL of ammonia gas at STP is compressed to 24 mL and 110 kPa. What must the new temperature of the gas be? Explain the result in terms of kinetic molecular theory.
2. I Design an investigation to test Dalton's law of partial pressure. Write a procedure and list the materials and equipment you need.
3. C Explain Dalton's law of partial pressure.
4. K/U What are the main components of the atmosphere?
5. MC Popcorn pops as the water in the kernel is vapourized. The building pressure becomes too much for the shell of the kernel, and it explodes. Explain, using your knowledge of the behaviour of gases, what happens to the temperature just before and just after the kernel pops. (Use a diagram to help you to visualize the situation.)

11.5 Gas Applications

Section Preview/ Specific Expectations

In this section, you will

- **identify** technological products based on compressed gas
- **describe** how a knowledge of gases is used in other areas of study
- **communicate** your understanding of the following term: *fuel cell*

In the previous section, you learned about the colourless, odorless mixture of gases that make up the atmosphere. It is easy to take the air around us for granted. You know that you need oxygen to breathe, but does it serve any other purposes? Nitrogen is not required for respiration, but it composes about four-fifths of our atmosphere. Is nitrogen a "useless" gas? In this section, you will learn about how gases are used. You will find out more about oxygen and nitrogen, and discover the importance of gases in deep-sea exploration.

Compressed Oxygen

You may never really think about breathing—until it becomes hard to do. In certain situations, normal human respiratory functions are disrupted. Hospitals use compressed oxygen for patients with respiratory disorders such as emphysema, pneumonia, or lung cancer.

You may have heard of hyperbaric oxygen (HBO) chambers being used to treat sports injuries. A high oxygen concentration in the blood causes blood vessels to constrict. This lowers the swelling that can occur in injured tissues. How is this high oxygen content delivered to the blood? In an HBO chamber such as that shown in Figure 11.28, air is compressed to three times normal atmospheric pressure. The patient breathes pure oxygen ($O_{2(g)}$) through a mask. The highly compressed air forces $O_{2(g)}$ in the lungs into the blood. This increases of dissolved $O_{2(g)}$ in the bloodstream. Consequently, oxygen is delivered rapidly to all the blood vessels of the body. In fact, oxygen can reach the injured tissue at 15 times the normal rate!

The use of compressed oxygen can benefit not only athletes, but also vulnerable premature babies. Premature babies can be afflicted with hyaline-membrane disease. This condition prevents the alveoli in their lungs from inflating, which leads to serious breathing difficulties. Placing these babies in an oxygen-rich environment such as an HBO or an incubator (Figure 11.28) helps inflate the alveoli. This increases the infants' chances for survival.

Figure 11.28 **Medical applications of pressurized oxygen can help the very strong, the very weak, and everyone in between.**

Oxygen at High Altitudes

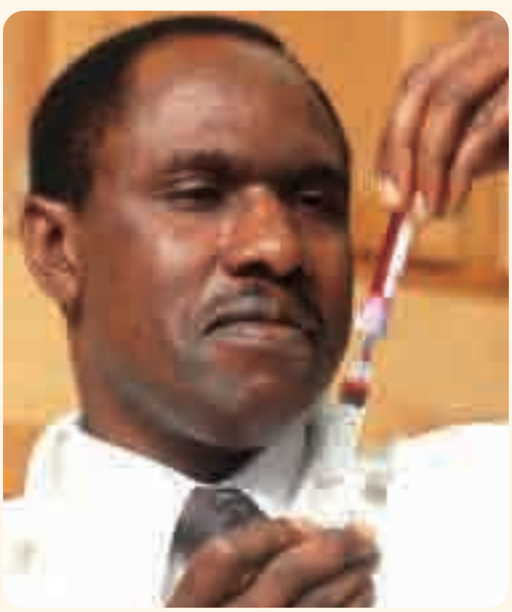
At other times, respiratory functions may be normal, but the environment poses problems. Commercial jet planes fly at an altitude of around 9.5 km. Airplanes carry pressurized oxygen for their passengers to breathe as they fly at high altitudes. Mountain climbers also need compressed oxygen when climbing to great heights. Those exploring ocean depths need oxygen as well. Scuba divers and submarine crews need compressed oxygen to breathe when they are underwater. You will learn more about undersea exploration later in this section.

Careers in Chemistry

Cancer Research Specialist

The National Cancer Institute of Canada estimates that in the year 2000 there were more than 132 000 new cases of cancer and 65 000 deaths caused by cancer in Canada. People are desperate for treatment.

Dr. Abdullah Kirumira, President and lead scientist at BioMedica Diagnostics in Windsor, Nova Scotia, has developed a way to speed up drug testing. This method uses compressed carbon dioxide and a dry chemistry diagnostics system including the firefly enzyme, luciferase.



Growing Cancer Cells for Testing

To obtain a consistent supply of cancer cells to test, CO_2 is pumped from a compressed gas cylinder. It is depressurized through a regulator into a specially designed incubator kept at about 37°C, the optimum temperature for cell growth. This CO_2-enriched atmosphere plays two roles:

- it is required for rapid cellular metabolism; and
- it interacts with a bicarbonate buffer in the growth medium to maintain an optimum pH balance.

Global Ties

In 1975, during the reign of dictator Idi Amin, Dr. Kirumira fled his native Uganda to Iraq, where he earned a B.Sc. in food technology. He then travelled to England for a Master's degree in dairy chemistry. From there, he moved to Australia where he earned his Ph.D. in biotechnology. In 1990, he and his wife moved to Nova Scotia in search of a slower-paced lifestyle.

At BioMedica, Dr. Kirumira oversees 12 employees, six of whom have advanced degrees, three technicians and three administrators. In addition, Dr. Kirumira teaches biochemistry and biotechnology at Acadia and Dalhousie Universities.

He frequently visits his home in Uganda. From Uganda, Dr. Kirumira says he can "take back what I have learnt from the West to help the people who probably need it the most."

Make Career Connections

1. To learn more about the biotech applications of using the firefly enzyme, luciferase, you can go to BioMedica's web site.
2. What universities offer degrees in biotechnology? Do research to find two universities near you that offer this program.

Web LINK

www.school.mcgrawhill.ca/resources/

CO_2 is a greenhouse gas, thought to contribute to global warming. Many countries around the world have sought to reduce their emissions of CO_2. International climate control treaties, such as the Montréal Protocol, have set targets to do this.

Since plants cycle CO_2 out of the atmosphere, several countries including Canada have argued that countries with large forest resources should be able to claim them as "carbon sinks." Rather than cutting their CO_2 emissions by the target amount, Canada and other countries maintain that they should be allowed to claim a reduction for the CO_2 absorbed by their large forests.

Research the current debate over "carbon sinks." Do you agree with the idea that CO_2 absorbed by plant life provides a valid strategy for controlling pollution? Go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next.

Oxygen and Combustion

A fire breaks out and smoke begins to fill the room. Breathing becomes difficult. Should you open a window to let the smoke out before you leave through the door? You may have learned you should never open windows before leaving the area of a fire. A fire can only burn as long as it has sufficient oxygen. (See Figure 11.29.)

Oxygen is an extremely important gas because it supports combustion. (You will learn more about combustion reactions in Unit 5.) The more oxygen available, the hotter a fire will burn. Pressurized oxygen is used in the manufacture of steel and specialized alloys. In the manufacture of steel, oxygen is used to remove excess carbon by burning the carbon into carbon monoxide or carbon dioxide. High levels of carbon make steel too brittle for many uses.

A remarkable invention called the **fuel cell** is used by the space industry. Some fuel cells burn hydrogen and oxygen gas, leaving water as a by-product (which astronauts can then drink). Companies such as Ballard Power Systems Inc. of Vancouver, British Columbia are working to make this technology practical for automobiles. Fuel cells are completely non-polluting. However, the current process for making large amounts of hydrogen gas does create a good deal of pollution, in the form of CO_2. Coal or another hydrocarbon is used to heat up and decompose water in a reaction similar to the following:

$$CH_{4(g)} + 2H_2O_{(g)} \xrightarrow{\text{heat}} CO_{2(g)} + 4H_{2(g)}$$

Before astronauts can use the fuel cell in space, they first have to get there! Liquid oxygen is an important component of rocket fuel. Liquid hydrogen and liquid oxygen are mixed and ignited in a combustion chamber. This reaction provides a very rapid expansion of gas. It produces enough energy to lift extremely heavy rockets carrying crewed vessels, such as the Space Shuttle, into orbit (Figure 11.30).

Figure 11.29 Liquid water becomes water vapour when exposed to the heat of a fire. Firefighters must take precautions to avoid steam burns.

Figure 11.30 The large centre tank that is attached to the shuttle orbiter has two compartments, one containing liquid hydrogen and the other containing liquid oxygen. The energy released by the reaction between these two substances provides the thrust needed to propel the orbiter into space.

Nitrogen's Many Uses

Inert gases are not very reactive. The word *inert* is also a synonym for sluggish and slow. Something with the property of inertness hardly sounds useful! Yet this very property makes nitrogen extremely valuable in technological and industrial applications.

Because nitrogen gas reacts with very few substances, it can be used to blanket substances, preventing them from reacting with oxygen. (A reaction with oxygen is called oxidation.) *Blanketing* means covering something with nitrogen, displacing all of the oxygen. This usually happens in a closed container. For example, ground coffee can taste bitter when exposed to oxygen for long periods of time. Packaging coffee in a nitrogen-enhanced container helps coffee keep its flavour. Gaseous nitrogen is also used for the long-term storage of fruits such as apples, allowing us to enjoy a wide variety of produce out of season.

Liquid nitrogen's low boiling point (77 K) means that it can be used in food preservation. Foods frozen quickly in liquid nitrogen retain more nutrients than slowly frozen foods. Less damage is done to cell structure during the quick-freezing process. Freezing foods this way also removes moisture, decreasing their weight and size. Freeze-dried foods are often used by people who require easily portable, lightweight food, such as campers, the military, and space explorers.

In cryosurgery, liquid nitrogen is used to fast-freeze cancerous tissues and warts. This proven new technique kills the cancerous area and allows surgeons to safely remove the dead tissue.

Gases and Undersea Exploration

Throughout history, humans have tried to explore the little-known world of the deep seas. Even with modern diving gear, the deep oceans are a dangerous place for humans. The air that a diver breathes to stay alive underwater can itself be one of the greatest hazards. That's because of the tremendous pressure exerted by the waters above a diver. For roughly every 10 metres of depth, water pressure on the diver's body adds the equivalent of one unit of atmospheric pressure (Figure 11.31).

The air that a diver breathes (mainly nitrogen and oxygen) must equal the pressure of the surrounding water if the diver's lungs are to stay inflated. Of course, this gas pressure increases as a diver goes deeper. As the gas pressure increases, the diver's bloodstream and body tissues absorb higher and higher volumes of gases.

Beyond depths of roughly 40–60 metres, the increased volumes of gas in the diver's body can cause *nitrogen narcosis*. When nitrogen narcosis strikes, normal feelings become dangerously exaggerated. A diver may experience a blissful giddiness, along with disorientation and impaired judgement. This condition is called "rapture of the deep." Divers in this state have been known to remove their breathing apparatus to swim like fish, mistake up for down, or simply lose track of how long they have been underwater.

At the other extreme, a diver's normal sense of caution may degenerate into irrational fear or even panic. A panicked diver may do the very

Figure 11.31 At 50 m underwater, a diver experiences pressure about five times higher than normal.

worst thing—rush to the surface without pausing to decompress on the way up. Why is this dangerous?

The dissolved gases such as nitrogen in the diver's bloodstream are under pressure. This is much like the carbon dioxide that is dissolved in a can of soda. If the diver returns to the surface too quickly, these dissolved gases behave similarly to the gas in soda when a can is opened. The gases escape from the blood as bubbles in the diver's blood vessels. These bubbles can block the flow of blood. They produce a painful and potentially fatal condition called "the bends."

Both nitrogen narcosis and the bends make prolonged deep diving a risky business. They limit both the safe depth and safe duration of human diving. Even skilled professional divers rarely descend beyond about 50 metres. Also, they rarely remain at that depth for much more than 30 minutes.

Section Wrap-up

Electronic Learning Partner

Your Chemistry 11 Electronic Learning Partner has a demonstration of CO_2 fire extinguishers and the chemistry behind their use.

In this section, you learned about some technological products and applications involving gases. You also learned that a knowledge of gases is important in other areas of study, such as medicine and deep-sea exploration.

In Chapter 12, you will find out about the ideal gas law. This law covers the many different gas laws you explored in this chapter. You will also discover a practical application for Dalton's law of partial pressures. You will learn how to do stoichiometric calculations for reactions that consume or produce a gas. In the laboratory, you will have a chance to produce and collect a gas. At the end of the next chapter, you will examine some of the chemistry that takes place in our atmosphere.

Section Review

Unit Issue Prep

What gases are produced as a result of human technologies such as fuel-burning engines?

1. C Why is fuel-cell technology, in its present state, not a pollution-free alternative to the internal combustion gasoline engine?
2. K/U Describe three industrial uses for liquid nitrogen.
3. MC How might covering apples with nitrogen or carbon dioxide gas preserve them longer through the winter?
4. K/U How does a hyperbaric oxygen chamber work to assist human healing?
5. MC Look up the standard mixture of gases for scuba diving tanks. What problem arises in using nitrogen gas in scuba tanks? How is this problem solved?
6. MC What does lightning (or bacteria) do to the nitrogen in the diatomic gas in order to make the atoms useful to plants? Use what you know about compounds and elements to answer this question.
7. MC Interview the science teachers at your school. Are there any compressed gas cylinders in the school? Ask the teachers about the appropriate safety precautions for compressed gases. Then prepare a short safety report on compressed gases in your school.

CHAPTER 11 Review

Reflecting on Chapter 11

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines.

- Describe the structure of solids, liquids, and gases in simple terms.
- Compare the intermolecular forces that hold solids together with those that hold liquids and gases together.
- Use the kinetic molecular theory to describe the behaviour of ideal gases.
- Explain Boyle's law.
- Explain Charles' law.
- Explain Gay-Lussac's law.
- Describe how gases under pressure are the cause of some of the natural wonders of our Earth.
- Describe how gases under pressure have many industrial and medical uses.
- Explain Dalton's law of partial pressure.

Reviewing Key Terms

For each of the following terms, write a sentence that shows your understanding of its meaning.

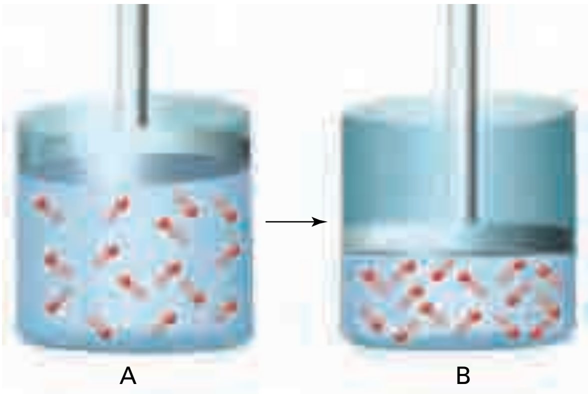
atmospheres	absolute zero
Boyle's law	Charles' law
closed system	combined gas law
Condensation	fuel cell
Dalton's law of partial pressures	Gay-Lussac's law
	kilopascals
fusible plugs	mmHg
ideal gas	pressure-relief valve
Kelvin scale	standard temperature and pressure (STP)
kinetic molecular theory of gases	standard temperature
pascal	standard atmospheric pressure
standard atmospheric pressure	
standard ambient temperature and pressure (SATP)	

Knowledge/Understanding

1. Gases behave differently than solids and liquids. Using what you know about forces between molecules, explain some of these differences.
2. Using the kinetic molecular theory, answer each of the following questions.
   (a) What does it mean when we say that the pressure of a gas has increased?
   (b) What does it mean when we say that the temperature of a gas has decreased?
3. (a) Give a real example that illustrates Boyle's law.
   (b) Give an example that illustrates Charles' law.
4. What is the kinetic molecular theory explaination for 0 K?
5. Explain Gay-Lussac's law in terms of the motion of gas particles.
6. Explain Dalton's law of partial pressures in terms of the motion of gas particles.
7. How is a volcano an example of a natural occurrence in which pressures of gases are of major importance?
8. List four major and four minor components of the atmosphere.
9. How is pure oxygen used medically and industrially?
10. Nitrogen is one of the most important industrial substances produced on Earth. Why is it so important industrially?
11. Would the column of mercury in a barometer be shorter at the top of a mountain or at the base of the mountain? Explain.
12. Using the kinetic molecular theory, explain why the density of a gas is less than that of a liquid. (Density is mass per unit of volume.)
13. A weather balloon containing helium is released into the atmosphere. As it rises, atmospheric pressure and temperature both decrease. Explain why the size of the balloon increases.

Inquiry

14. Below are two diagrams showing a trapped gas. The volume of the gas in container A can be changed to become the same volume as the gas in container B. Describe how you could do this in the laboratory.



15. You have carried out investigations to determine the relationship between pressure and volume, and between temperature and volume. How could you investigate the relationship between pressure and temperature?
 (a) Write up a procedure to determine the relationship between the pressure and temperature of a gas.
 (b) Identify any problems with your procedure. How could you overcome these problems?
 (c) What materials will you need?
 (d) What variables will you hold constant?
 (e) What variable will you change?
 (f) What variable will you measure?
16. The pressure exerted on 0.25 L of nitrogen is 120 kPa. What volume will this gas occupy at 60 kPa if the temperature and number of moles is constant?
17. Ammonia gas, $NH_{3(g)}$, is used in the production of fertilizer. At 55.0˚C, a sample of ammonia gas is found to exert a pressure of 7.5 atm. What pressure will the gas exert if its volume is reduced to one fifth of its original volume at 55.0˚C?
18. A 35.0 L sample of dry air at 750 torr is compressed to a volume of 20 L. What is the final pressure exerted on the gas if the temperature remains constant?
19. A 25 g sample of dry air in a large party balloon at 20˚C occupies a volume of 20 L. If the temperature is increased to 40˚C at constant pressure, how large will the balloon become?
20. A sample of nitrogen gas has a volume of 10 L at 101.3 kPa and 20˚C. To preserve biological tissue, the nitrogen gas is cooled to –190˚C, almost the temperature of liquid nitrogen, at a pressure of 101.3 kPa. What volume will the nitrogen occupy at this temperature?
21. A 75.3 L sample of oxygen gas at 25.7˚C is cooled until its final volume becomes 10 L. If the pressure remains constant, what is the final temperature (in ˚C)?
22. Calculate the volumes that each of the following gases would occupy at STP.
 (a) 22.4 L of oxygen at 75˚C and 700 torr.
 (b) 100 cm^3 of nitrogen at 20˚C and 150.0 kPa.
 (c) 45 mL of neon at –50.˚C and 200 kPa.
23. A birthday balloon contains 2.0 L of air at STP. What volume will the balloon have at SATP?
24. A mixture of neon and argon gases is collected at 102.7 kPa. If the partial pressure of the neon is 52.5 kPa, what is the partial pressure of argon?
25. A 250 mL glass vessel is filled with krypton gas at a pressure of 700 torr at 25.0˚C. If the glass vessel is made to withstand a pressure of 2.0 atm, to what maximum temperature (in ˚C) can you safely heat the flask?
26. A cylinder with a moveable piston contains hydrogen gas collected at 30˚C. The piston is moved until the volume of the hydrogen is halved. The pressure inside the container has increased to 125 kPa at 30˚C. What was the initial pressure inside the cylinder?
27. Argon gas is used inside light bulbs because it is a plentiful inert gas. 650 cm^3 of argon gas at STP is heated in order to double its volume at 101.3 kPa. What is the final temperature (in ˚C)?
28. A truck leaves Yellowknife in early January when the temperature is –30.0˚C. The tires of the truck are inflated to 210 kPa. Four days later, the truck arrives in California where the temperature is 30.0˚C. What is the air pressure in the tires when the truck arrives at its destination?
29. Methane (CH_4), a natural gas, is stored in a 100 L tank at –10˚C and a pressure of 125 atm. The gas is used to provide fuel for a furnace that heats a country home during the winter. The furnace consumes an average of 500 L of methane a day. How long will this supply of methane last if it is burned at 450˚C at a pressure of 102 kPa?

30. Neon gas is widely used as the luminous gas in signs. A sample of neon has a volume of 5.5 L at 750 torr at 10.0˚C. If the gas is expanded to a volume of 7.5 L at a pressure of 400 torr, what will its final temperature be (in ˚C)?

31. Halogen lamp bulbs are usually filled with bromine or iodine vapour at 5.0 atm pressure. When turned on, the glass bulb can heat up to more than 1150˚C. If room temperature is 20˚C, what will the pressure in the bulb be when it reaches its operating temperature?

32. Helium gas is stored in a steel cylinder with a volume of 100 L at 20˚C. The pressure gauge on the cylinder indicates a pressure of 25 atm. The cylinder is used to blow up a weather balloon at 25˚C. If the final pressure in the cylinder and the balloon is 1.05 atm, how large will the balloon be?

33. A scuba diver is swimming 30.0 m below the surface of Lake Ontario. At this depth, the pressure of the water is 4.0 atm and the temperature is 8.0˚C. A bubble of air with a volume of 5.0 mL escapes from the diver's mask. What will the volume of the bubble be when it breaks the surface of the water? The atmospheric pressure is 101.3 kPa and the temperature of the water is 24.0˚C.

Communication

34. The weather person on television reports that the barometric pressure is 100.2 kPa. How high a column of mercury will this air pressure support?

35. Convert each of the following pressures:
   (a) 1.5 atm to kPa
   (b) 135.5 kPa to mm Hg
   (c) 750 mm Hg to torr

36. On a visit to your doctor, your blood pressure is taken. The reading is 125.0 mm Hg systolic and 80.0 mm Hg diastolic. What is your blood pressure in kPa?

37. Convert the following temperatures.
   (a) –185.5˚C to K
   (b) 125 K to ˚C

38. Do research to find out how carbonated drinks get their "fizz." Prepare a short PowerPoint™ presentation to the class that describes the manufacture of a typical soft drink.

Making Connections

39. Coal is a fossil fuel that is burned to obtain energy. However, carbon dioxide (CO_2) gas is produced whenever coal is burned. What are the benefits and risks of burning coal to obtain energy? What are the alternatives?
   (a) Do research to answer these questions.
   (b) *Should the government shut down all coal-burning power plants in Canada to help prevent global warming?* Carry out this debate with your class. Divide the class into the executives of a coal-burning power plant, and scientists representing an environmentalist group.

40. Interview a doctor or nurse who works at a hospital. Find out how compressed gases are used at the hospital, and what safety precautions are taken.

41. Suppose you are a member of a consulting firm that is approached by the Ministry of the Environment. You are asked to prepare a report on how the government can help reduce CO_2 and CH_4 gas emissions. Both gases are important greenhouse gases that may be causing global warming.
   (a) Identify the information you will need to prepare this report. This may take the form of a list of questions.
   (b) Research the answers to your questions.
   (c) Prepare a short report giving suggestions on how the Canadian government could help reduce CO_2 and CH_4 gas emissions.

Answers to Practice Problems and Numerical Section Review Questions

Practice Problems: **1.** 1.0×10^3 kPa **2.** 1644 L **3.** 4.5 cm^3 **4.** 0.01 L **5.(a)** 298 K **(b)** 310 K **(c)** 423 K **6.(a)** 100˚C **(b)** –175˚C **(c)** 152˚C **8.** 3.86×10^2 mL **9.** 2.1 L **10.** 606˚C **11.** 957˚C **12.** 127 L **13.** 2.71×10^2 kPa **14.** 18.8 atm **15.** –16˚C **16.** 31˚C **17.** 163 mL **18.** 0.43 atm **19.** 148 K **20.** 310˚C **21.** 6.4×10^2 kPa **22.** 300.6 kPa **23.** 14 kPa, 27 kPa, 75 kPa **24.** 1.08×10^2 kPa **25.** 0.70 atm
Section Review: **11.1** **3.(a)** liquid **(b)** gas **(c)** solid
11.2 **1.(a)** 206 kPa **(b)** 0.841 atm **(c)** 1.14×10^3 torr **(d)** 80.0 kPa **4.** 1.04 L **5.** 152 kPa **6.** 811 torr
11.3 **2.(a)** 52.1 kPa **(b)** 8.3 L **11.4** **1.** 148 K