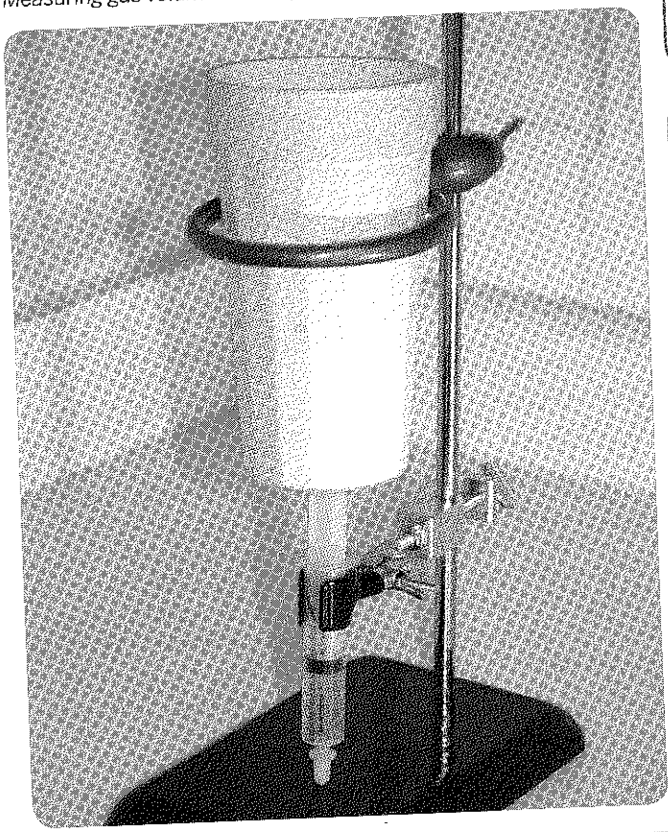
LABORATORY 14.1:

OBSERVE THE VOLUME-PRESSURE RELATIONSHIP OF GASES (Boyle's Law)

Boyle's Law states that, at constant temperature, the volume of a gas is inversely proportional to its pressure. In this laboratory session, we'll verify Boyle's Law experimentally, using the setup shown in Figure 14-1. (You can buy a ready-made Boyle's Law apparatus from Home Science Tools or other vendors, but $10 or $12 is a pretty high price to pay for a disposable syringe and two blocks of wood.) The apparatus shown in Figure 14-1 uses only standard lab equipment, and is at least as accurate as a ready-made apparatus.

FIGURE 14-1:
Measuring gas volume as compressed by a measured mass



REQUIRED EQUIPMENT AND SUPPLIES

- ☐ goggles, gloves, and protective clothing
- ☐ balance and weighing cup
- ☐ caliper
- ☐ barometer (optional)
- ☐ ring stand
- ☐ burette or utility clamp (to fit syringe)
- ☐ 4" (100 mm) support ring
- ☐ plastic syringe, 10 mL to 50 mL, graduated, with cap
- ☐ mineral oil or petroleum jelly (1 drop)
- ☐ plastic cup (to fit support ring)
- ☐ lead shot (10 pounds or 5 kilograms)

SUBSTITUTIONS AND MODIFICATIONS

- If you do not have a caliper, you may substitute a metric ruler with millimeter markings, although you will sacrifice significant accuracy.
- If you do not have a barometer, you may use the barometric pressure broadcast by a local TV or radio station or the Weather Channel web site for your zip code, but see the note on the next page.
- You may substitute any sturdy lightweight container of similar size for the plastic cup, including an aluminum beverage can with the top removed. The container should fit loosely within the support ring: not closely enough to bind, but closely enough to keep the container centered over the syringe plunger when mass is added to the container.
- You may substitute any dense material for the lead shot, such as old wheel weights, fishing sinkers, spools of solder, and so on. Ideally, the material should be dense enough to allow the container to hold at least 3 kilograms of mass. Syringes with large bores require more mass for equivalent compression.

Using this apparatus, we'll begin with the syringe containing a known volume of gas under normal atmospheric pressure. We'll then add mass incrementally to the container above the syringe to increase the pressure on the gas contained in the syringe and record the gas volume under differing amounts of pressure.

Gas pressure is specified in units of mass, weight, or force per unit area. For example, in traditional units, standard atmospheric pressure is about 14.7 pounds per square inch. Chemists use the SI unit of pressure, the ***pascal (Pa)***, which equals one newton per square metre ($N \cdot m^{-2}$ or $kg \cdot m^{-1} \cdot s^{-2}$). In SI units, standard atmospheric pressure is about 101,325 Pa, which may also be stated as 101.325 kilopascal.

The total pressure exerted on the gas in the syringe is the sum of the atmospheric pressure, the pressure exerted by the mass of the syringe plunger and the container above it, and the pressure exerted by the mass added to the container. But there's one more piece of the puzzle. Because pressure is specified per unit area, we need to know the area of the syringe bore. With all of those data known, we can calculate the pressure of the gas contained within the syringe and correlate that pressure with the observed volume of the gas.

ATMOSPHERIC PRESSURE VERSUS BAROMETRIC PRESSURE

Although most people assume that ***atmospheric pressure*** and ***barometric pressure*** mean the same thing, they don't. Atmospheric pressure is the actual pressure of the atmosphere. Barometric pressure, the value given by TV and radio stations (and indicated by your barometer if you adjusted it to the pressure given in a local weather report), is an adjusted value that reflects the elevation of the reporting station. Barometric pressure is what the atmospheric pressure would be *if the reporting station were located at sea level*. For a reporting station located at sea level, barometric pressure is equal to atmospheric pressure. But for a reporting station located above sea level, the actual atmospheric pressure is lower—sometimes significantly lower—than the barometric pressure reported by the station.

To obtain an accurate atmospheric pressure reading, you can adjust the barometric pressure reported by a station for that station's altitude. Alternatively, you can check with your local airport, which reports both barometric pressure and atmospheric pressure.

LABORATORY 14.2:

OBSERVE THE VOLUME-TEMPERATURE RELATIONSHIP OF GASES (Charles' Law)

Charles' Law states that, at constant pressure, the volume of a gas is proportional to its absolute temperature, specified in kelvins. According to Charles' Law, then, doubling the temperature of a gas—say, from 200 K to 400 K—doubles its volume, and halving the temperature halves the volume. This proportionality holds true regardless of the percentage change. For example, if we increase the temperature of a 7.5 mL gas sample from 293.15 K (20.00°C) to 373.15 K (100.00°C), we can calculate the volume at the higher temperature by substituting the known values in the equation for Charles' Law:

$$V_1 \cdot T_2 = V_2 \cdot T_1$$

or

$$(7.5 \text{ mL}) \cdot (373.15 \text{ K}) = (x \text{ mL}) \cdot (293.15 \text{ K})$$

Solving for x, we find that the gas volume at the higher temperature is about 9.5 mL.

Similarly, if we decrease the temperature of a 7.5 mL gas sample from 293.15 K (20.00°C) to 194.65 K (−78.50°C), we can calculate the volume at the lower temperature, again by substituting the known values in the Charles' Law equation:

$$(7.5 \text{ mL}) \cdot (194.65 \text{ K}) = (x \text{ mL}) \cdot (293.15 \text{ K})$$

Again solving for x, we find that the gas volume at the lower temperature is about 5.0 mL.

I didn't choose that volume and those temperatures arbitrarily. The syringe I used in my apparatus has a full-scale reading of 10.0 mL, so an initial volume of 7.5 mL at 20°C (about room temperature) allowed me to get near (but not exceed) the full-scale 10.0 mL graduation when heating the syringe to 100°C. That's the approximate temperature of boiling water, which was the highest temperature I was comfortable using for this experiment. I could have substituted vegetable oil or another higher-boiling liquid and gotten up to 150°C—the upper limit of my thermometer's scale—but doing that would add little to the

REQUIRED EQUIPMENT AND SUPPLIES

- ☐ goggles, gloves, and protective clothing
- ☐ thermometer
- ☐ beaker, 150 mL or larger (2)
- ☐ ring stand
- ☐ burette or utility clamp (to fit syringe)
- ☐ 4" (100 mm) support ring
- ☐ wire gauze
- ☐ alcohol lamp, gas burner, or other heat source
- ☐ plastic syringe, 10 mL to 50 mL, graduated, with cap
- ☐ mineral oil or petroleum jelly (1 drop)
- ☐ ethanol, isopropanol, or acetone (sufficient to fill beaker; needed only if you use dry ice)
- ☐ ice
- ☐ dry ice (optional, see Substitutions and Modifications)

WHAT'S A KELVIN?

Note that the SI unit of temperature is not "degrees kelvin" or "kelvin degrees" but just "kelvins" all by itself. The Kelvin temperature scale uses the same incremental units of temperature as the familiar Celsius scale. In other words, increasing or decreasing the temperature by one degree Celsius (1°C) also increases or reduces the temperature by one kelvin (1 K).

The difference between the Kelvin scale and the Celsius scale is the baselines. The Kelvin scale assigns a temperature of 0 K to absolute zero, which is the coldest possible temperature, where even atomic vibrations cease. The Celsius scale assigns a temperature of 0°C to the melting point of pure water, which can also be specified as 273.15 K. It follows, therefore, that on the Celsius scale, absolute zero is −273.15°C.

SUBSTITUTIONS AND MODIFICATIONS

- If you use dry ice, the thermometer needs to be accurate down to at least −80°C. If you use a standard freezer as your cold bath, the thermometer needs to be accurate down to −20°C or so. Typical inexpensive digital thermometers are accurate down to −40 or −50°C, and typical inexpensive glass thermometers are accurate down to −20°C.
- If you do not have a ring stand, clamp, and so on, you may substitute beaker tongs and a large steel bolt or other weight. Tie the bolt securely to the body of the syringe (not the plunger) and use the weight to keep the syringe fully immersed in the hot and cold baths. Use the beaker tongs to add and remove the syringe from the baths.

value of the experiment, increase the danger level, and use up a lot of vegetable oil for no good reason.

I wanted the temperature of my cold bath as low as possible. In my first pass, I put the apparatus in a full-size freezer, which got the temperature down to about −20°C. I was able to do a bit better than that, though. As I was working on this chapter, I happened to receive a FedEx shipment of food packed in dry ice. Putting chunks of dry ice in a beaker full of alcohol allowed us to reach a temperature considerably colder than our freezer. You can purchase small amounts of dry ice locally. Check the Yellow Pages for "dry ice."

In this laboratory session, we'll use an apparatus similar to the one we used in the preceding lab. The difference is that in the previous lab we held temperature constant and changed the pressure, and in this lab we hold pressure constant and change the temperature. Thus, we won't need the container and the masses we used in the preceding lab. Instead, we'll record the volume contained by the syringe at room temperature, and then immerse the syringe in liquids at various temperatures and record the changes in volume.

CAUTIONS

Use extreme care with dry ice. At atmospheric pressure, dry ice sublimates (changes directly from a solid to a gas) at −78.5°C, cold enough to cause severe frostbite almost instantly if it contacts your skin. Handle dry ice only with tongs. Make sure the alcohol lamp or gas burner is widely separated from the beaker of alcohol—across the room is best. Wear splash goggles, heavy-duty gloves, and protective clothing.

LABORATORY 14.3:
OBSERVE THE PRESSURE-TEMPERATURE RELATIONSHIP OF GASES (Gay-Lussac's Law)

Gay-Lussac's Law states that, at constant volume, the pressure of a gas is proportional to its absolute temperature, specified in kelvins. For example, if you double the temperature of a gas, you double its pressure, and vice versa. Gay-Lussac's Law can be expressed as the equation:

$$P_1 \cdot T_2 = P_2 \cdot T_1$$

For example, if we increase the temperature of a 7.5 mL gas sample at atmospheric pressure from 293.15 K (20.00°C) to 373.15 K (100.00°C), we can calculate the pressure at the higher temperature by substituting the known values in the Gay-Lussac's Law equation:

$$(101{,}325 \text{ Pa}) \cdot (373.15 \text{ K}) = (x \text{ Pa}) \cdot (293.15 \text{ K})$$

Solving for x, we find that the gas pressure at the higher temperature is about 128,976 Pa.

In this lab, we'll verify Gay-Lussac's Law experimentally by using the apparatus shown in Figure 14-4, along with some of the measurements and data we recorded in the first lab session in this chapter. We'll start with a known volume of air in the syringe, with the syringe in a water bath at room temperature. We'll then heat the water bath to boiling, which will cause the volume of air in the syringe to increase. With the gas sample held at constant temperature in the boiling water bath, we'll add mass to the container until the syringe is depressed to its original 7.5 mL volume reading. Knowing the mass required to compress the hotter gas sample to its original volume, we can calculate the pressure of the 7.5 mL gas sample at that higher temperature.

SUBSTITUTIONS AND MODIFICATIONS

- If you are using the same syringe, container, and other components you used in Laboratory 14.1, you may use the measurements and calculations you did for that laboratory session rather than repeating them here.
- Other substitutions and modifications are as listed in Laboratory 14.1.

REQUIRED EQUIPMENT AND SUPPLIES

- ☐ goggles, gloves, and protective clothing
- ☐ balance and weighing cup
- ☐ caliper
- ☐ barometer (optional)
- ☐ ring stand
- ☐ burette or utility clamp (to fit syringe)
- ☐ 4" (100 mm) support ring (2)
- ☐ wire gauze
- ☐ alcohol lamp, gas burner, or other heat source
- ☐ beaker, 150 mL
- ☐ plastic syringe, 10 mL to 50 mL, graduated, with cap
- ☐ mineral oil or petroleum jelly (1 drop)
- ☐ plastic cup (to fit support ring)
- ☐ lead shot (10 pounds or 5 kilograms)

CAUTIONS

The real hazard in this lab is that the apparatus may topple or collapse as you add too much mass to the container, splashing boiling water everywhere. Take extreme care with the hot water, make sure that the container remains centered over the syringe plunger as you add mass, and use the smallest beaker that allows the gas-filled portion of the syringe to be fully immersed. If you have a third support ring, use two rings to surround the mass container to prevent it from tipping. Wear splash goggles, gloves, and protective clothing.

FIGURE 14-4: *Our apparatus for verifying Gay-Lussac's Law*