Activity 5

Gas Pressure

GOALS
In this activity you will:

• Generate CO₂ by reacting an effervescent antacid tablet with water.
• Examine the relationship between temperature and pressure produced by an enclosed gas.
• Relate Charles’s Law and Gay-Lussac’s Law with the Kinetic Molecular Theory.

What Do You Think?

After putting water and an effervescent antacid tablet into a film canister, gas is created and you can expect the film canister to “flip its lid.”

• Why does the canister “flip its lid?”
• How could you adapt the process so that the top would not blow off?

Record your ideas about these questions in your Active Chemistry log. Be prepared to discuss your responses with your small group and the class.

Investigate

As you have learned in Activity 3, gases are made of tiny molecules that are in constant motion. These particles are constantly colliding and creating gas pressure inside the container holding the gas. How are the temperature and pressure of the gas related?

1. Add five quarters to an empty film canister. Add 6 mL of water. The quarters serve the purpose of preventing the film canister from tipping over and floating. They do not participate in the chemical reaction.
2. Grind up some effervescent antacid tablets in a mortar with a pestle. Weigh out some of the powder and then add it to the film canister.
   a) Record the mass in your Active Chemistry log.

3. Quickly snap on the lid, place it on the table, and step back.
   a) Record how long it takes before the lid pops off.

4. Continue to alter the amount of powder until the lid remains on the canister for a period of between 25 and 35 s (seconds) before popping. That means that it can’t pop before 25 s, but it will pop before 35 s. Examine the contents left in the film canister. There should not be any white solid remaining. If there is, reduce the amount of effervescent antacid tablets and try again. (You must empty out the contents of the canister and wipe it dry before repeating the experiment.) You might choose to use a table similar to the one shown below to keep track of your data.

5. Using the same canister, number of quarters, and volume of water as in Step 4, place the canister in a flat-bottom dish. Pour cold water in the dish to surround the canister to within 1 cm of its rim. You might want to add some ice to cool the water down to a temperature at least 15°C lower than the room temperature.

6. Allow the water and quarters to cool down to the temperature of the water in the dish.
   a) Measure and record the temperature of the water in the dish.

7. Add the same amount of powdered effervescent antacid tablets to the canister as determined in Step 4. Quickly snap on the lid, place it back in the cold water, and step back.
   a) Measure and record the time until the lid pops off.
   b) If the lid does not pop off, what must this tell you about the pressure of the gas inside the film canister at this lower temperature? Check to make sure no white solid remains.
   c) In your Active Chemistry log, analyze what you have seen and formulate the relationship between the pressure of a constant volume of gas and its temperature. State your evidence. How could you test this relationship?

8. Dispose of the materials as directed by your teacher. Clean up your workstation.

<table>
<thead>
<tr>
<th>Trial</th>
<th>Mass of Antacid Tablets (mg)</th>
<th>Time it Takes Top to Pop (s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
GAS LAWS

Gay-Lussac’s Law

Carbon dioxide is a very important and common chemical that is found in the stomach. Many antacids neutralize excess stomach acid and at the same time produce CO₂. All carbonated beverages have large amounts of CO₂ dissolved under pressure in the water. Some of this gas escapes when the bottle or can is opened, but a lot of gas still remains dissolved, which can be released upon warming. This is what makes a cold carbonated drink fizzy or effervescent on a hot day. Everyone has probably tasted a soft drink that has been left open and has been warmed. Most of the CO₂ has escaped and so the soft drink tastes flat.

When a large amount of gas is ingested or generated in the stomach quickly, there is a certain amount of discomfort. When there is too much gas in the stomach, it comes up through the esophagus and out the mouth as a burp. There is a specialized muscle at the junction of the esophagus and the stomach that remains closed except to allow food to pass from the esophagus to the stomach. Just as in the activity with the film canister, where the pressure rose until the cap popped, the gas pressure in your stomach rises until it overcomes the pressure of the muscle that is acting like a valve. The gas that rushes up the esophagus and out your mouth results in a burp. The warming of a carbonated drink in the stomach makes the CO₂ less soluble, releasing considerable gas. This almost always makes you burp after quickly drinking a cold soda. It is often mentioned that in some countries and cultures (probably not the USA) a burp after eating is a sign that the meal was very well prepared, and is a compliment to the cook!

Another important relationship that is useful when working with gases has been demonstrated in this activity. Just as in Activity 3 where the relationship between volume and temperature was referred to as Charles’s Law, the relationship between pressure and temperature is often referred to as Gay-Lussac’s Law. Yes, it is the same Joseph Louis Gay-Lussac associated with Charles’s Law. From the behavior of the CO₂ in the film canister, it is easy to see the relationship expressed in Gay-Lussac’s Law. The temperature and the pressure are directly proportional, if the mass and volume of the gas remain constant.

Remember, a direct proportion means that as one quantity goes down the other one goes down proportionately; and if one goes up, the other
goes up proportionately. Don’t forget, all temperatures must be absolute temperatures. In other words, all temperatures must be changed from Celsius temperatures into absolute Kelvin temperatures. To change from Celsius to absolute Kelvin temperatures add 273 to the Celsius temperature. When the temperature of the gas trapped in the film canister went down, so did the pressure. In fact, the pressure went down enough so that the cap was not blown off.

Gay-Lussac’s Law is the basis for the warning on aerosol cans that says “do not incinerate.” Even though the can appears empty, it retains a reasonable amount of gas. Exposing an aerosol can to extreme heat would be very dangerous because it will raise the temperature of the gas. Because the gas is trapped in a constant volume, it will explode as the pressure rises beyond the strength of the can. In doing so, it could splatter dangerously hot and/or flammable material all around.

**Kinetic Molecular Theory of Gases**

In Activity 3 and in this activity you have seen examples of Charles’s Law and Gay-Lussac’s Law. There are many other situations where these relationships among volume, temperature, and pressure hold. These laws led to the development of a theory that attempted to explain why these relationships were observed. It is known as the **Kinetic Molecular Theory of Gases**. The assumptions of this theory apply to all gases, but the equations derived from the theory are only accurate for what is known as an **ideal gas**. To understand what this theory says you must think small, at the molecular (nanoscopic) level. The fundamental assumptions that the scientists agreed upon in formulating this theory are:

1. All gases are made up of tiny particles. These particles are called atoms or molecules.
2. The particles in gases are very far apart in comparison to their own diameter. In other words, a gas is made up of mostly empty space.
3. These particles are in constant motion and move in straight paths until they collide with something.
4. When the particles collide they bounce off each other. Energy may be transferred to whatever they run into, but no energy is lost.
5. At any particular instant, all of the particles do not possess the exact same amount of energy.
However, the average kinetic energy of all the particles is directly proportional to their absolute (or Kelvin) temperature.

The equations that relate the pressure, volume, and absolute temperature of gases under different conditions are accurate only for an ideal gas (one whose molecules exert no attraction whatever for each other). However, as long as the temperature and pressure of a gas is close to room temperature and ordinary atmospheric pressure, these ideal gas equations can be used to describe real gases as well.

**Pressure**

Now, try to think what is happening at the molecular level. The gas **pressure** is due to the collisions that the particles have with the walls of the container. Two major factors should be considered in terms of the pressure that the gas exerts. The first is how often the particles collide with the walls, and the second is with what force they collide with the walls. The gas pressure inside a balloon will be greater if either the number of collisions in one second increases or if the molecules hit the inside walls with more forceful collisions. If you could be shrunk down and put inside a balloon, the beating that you would take from the gas molecules inside would depend on how often they were hitting you as well as how hard they hit you.

**Temperature**

**Temperature** is measured with a thermometer. Although your sense of touch gives you some indication of temperature, it does not have the precision of a thermometer. For example, the tile floor and the carpet in a bathroom will probably have the same temperature, but the tile floor may feel much colder because of how well it conducts heat energy. When working with gases, you find that temperature is a measure of the average kinetic energy of the particles in the gas. Some particles are traveling very fast and have large kinetic energies. They may collide with other slower-moving particles in the same gas and lose some of that kinetic energy. The average kinetic energy of all the particles is the temperature of the gas.

**Applying Charles’s Law**

Charles’s Law states that volume and temperature are directly proportional when the pressure of a constant mass of gas is held constant (refer back to Activity 3).
Imagine particles of gas inside a box with a moveable weighted cover, as shown in the diagram. The pressure of the gas is equal to the weight of the cover. What would happen if someone put the box in an oven?

- The temperature of the container and the gas inside would increase.
- The velocity of the particles of the gas and the kinetic energy would increase.
- The particles would be hitting the walls and cover with more force and more often. The pressure of the gas would increase.

If the pressure of the gas were to increase, the weight would move up until the pressure of the gas once again equaled the weight of the cover.

You can now compare the situation with the cold air and the hot air. The pressure is the same and is equal to the weight of the cover. The cool air requires a small volume and the hot air requires a large volume.

Charles’s Law can be used to determine how much the volume changes as a result of a change in temperature. If the old volume is divided by the old temperature, it will equal the new volume divided by the new temperature. This is one mathematical way of expressing a direct proportion.

\[
\frac{V_{\text{old}}}{T_{\text{old}}} = \frac{V_{\text{new}}}{T_{\text{new}}}
\]

For example, a balloon has a volume of 240.0 mL at the beach where the atmospheric pressure is 760 torr and the absolute temperature is 298 K. The pressures on the inside and outside of the balloon are equal, so the gas pressure is also 760 torr. When the Sun came out in the afternoon, the volume of the balloon increased to 245 mL.

What was the temperature in the afternoon? The atmospheric pressure was still 760.0 torr:

\[
\frac{240.0 \text{ mL}}{298 \text{ K}} = \frac{245 \text{ mL}}{T_{\text{new}}}
\]

Solving for \(T_{\text{new}}\) gives a temperature of 304 K.
Now check to make sure the answer seems reasonable. The pressure remained constant, so if the temperature goes up, the volume must also; and so it does. Always check to see if your answer seems reasonable! 304 K is 31°C (88°F) which is a reasonable value. This analysis of the balloon at the beach assumes that the extra stretch of the balloon does not add an additional pressure. Taking into account the pressure of the balloon material would make the problem more difficult but more accurate.

**Applying Gay-Lussac's Law**

Gay-Lussac’s Law says that pressure and temperature are directly proportional if volume is held constant. Imagine yourself shrunk down again to molecule size but trapped in a 55-gallon oil drum. All the oil has been removed and your only companions are the many, many particles that make up the air filling the drum. This time the volume cannot change. What would happen if somebody put a blowtorch to the outside of the drum?

- The particles would start moving faster. This means that they would be hitting the walls with force.
- Those fast-moving particles will also be hitting you and the wall more often.
- Because the volume can’t change, Gay-Lussac’s Law seems to make sense when it predicts pressure will also increase as the temperature increases.

If the drum rolled into the river, the temperature of the gas would decrease. This means the speed of the particles would decrease. And the number of collisions on you would also decrease.

Gay-Lussac’s Law can be used to find how much the pressure changes as a result of the change in the temperature. If the old pressure is divided by the old temperature, it will equal the new pressure divided by the new temperature.

\[
\frac{P_{\text{old}}}{T_{\text{old}}} = \frac{P_{\text{new}}}{T_{\text{new}}}
\]

For example, an aerosol can having a volume of 450 mL and a temperature of 292 K and a pressure of 950.0 torr was left out in the bright Sun and the temperature of the gas inside rose to 310 K. What was the resulting pressure inside the can?

\[
\frac{950.0 \text{ torr}}{292 \text{ K}} = \frac{P_{\text{new}}}{310 \text{ K}}
\]

Solving for \(P_{\text{new}}\) gives a pressure of 1010 torr (or 1.3 atm).
What Do You Think Now?

At the beginning of this activity you were asked:

- Why does a canister “flip its lid?”
- How could you adapt the process so that the top would not blow off?

How would you describe what the pressure of a gas is?

How would you describe what the temperature of a gas is?

Using a constant volume of gas, how are the behaviors of the temperature of the gas and the pressure of the gas related? Is it possible to change one without changing the other?

Chem Essential Questions

What does it mean?

Chemistry explains a macroscopic phenomenon (what you observe) with a description of what happens at the nanoscopic level (atoms and molecules) using symbolic structures as a way to communicate. Complete the chart below in your Active Chemistry log.

<table>
<thead>
<tr>
<th>MACRO</th>
<th>NANO</th>
<th>SYMBOLIC</th>
</tr>
</thead>
<tbody>
<tr>
<td>Describe how the pressure inside a film canister changed when temperature increased.</td>
<td>How do the moving gas particles in a canister exert pressure? Explain using the Kinetic Molecular Theory of Gases.</td>
<td>Using the gas law equations, show how they can be used to predict direct and indirect proportions.</td>
</tr>
</tbody>
</table>

How do you know?

What evidence do you have that the pressure of a gas increases with an increase in the temperature?

Why do you believe?

What happens to the pressure inside an automobile tire when the winter arrives and the temperature drops?

Why should you care?

Propose a miraculous escape from a section of the alimentary canal through using an apparatus that operates on a gas-law principle.
Reflecting on the Activity and the Challenge

In this activity, you learned about CO₂, one of the most important gases associated with your body. It also allowed you to learn about some of the relationships involving the pressure, temperature, and volume of gases as observed in the macroscopic world.

Now, try to put yourself into the nano-world of individual molecules. This would be a world where you would be shrunk down to one-billionth of your current size. Try to understand these relationships from this new perspective. This should enable you to better envision what the environment in the alimentary canal might be like. In this nano-world, the release of CO₂ from a soda would be a catastrophic event, greater even than a hurricane here in the macroscopic world.

Chem to Go

1. What happens to the pressure inside a sealed rigid container when the temperature of the container and gas increase?

2. What is the cause of a burp after drinking a cold soft drink?

3. What happens to the solubility of carbon dioxide gas as the temperature increases?

4. In your experiment involving the effervescent antacid tablet and water in the film canister, the cap blew off quickly in the first experiment and there was white solid powder left. The next time you used less antacid until finally you reduced the amount of material so that there was no white solid left when the reaction was over. What would happen if you did one more experiment and used even less antacid?

5. At room temperature, which is 300 K, an empty aerosol can contained 200 mL of gas under a pressure of 1.5 atm. The highest pressure that the can’s wall will withstand before exploding is 2.5 atm. If the can is exposed to extreme heat, at what temperature will it explode?

6. Air is pumped into a tractor tire at a pressure of 1500 torr. At dawn the temperature of the air in the tire is 295 K. At 3:00 PM in the afternoon the temperature is 308 K. What is the pressure in the tire now? (Assume that the volume of the tire did not change.)
7. Preparing for the Chapter Challenge

You know that the stomach is a very violent rocking and rolling environment when food is undergoing digestion. If your host body just guzzled a whole bottle of soda and it had made its way down the esophagus and was just entering the stomach and you were in the area, you might be in danger of being burped up and out through the mouth. It is very hard to match the conditions you would find in the stomach at this point, but it brings up an interesting problem. Do you think that you could “burp” off the lid of the film canister by pouring an ice-cold soda into a canister with a temperature of 37°C? You know it works in the stomach, but will it work in the film canister? See if you can devise a way to “burp” off the lid.

Inquiring Further

1. Aerosol cans

Read the label on two aerosol cans concerning safety and cautions with heat. What temperature is mentioned on the cans? What happens if you exceed that temperature?

2. Oxygen in water

Dissolved oxygen in a lake or pond is important for fish to survive. What happens to the solubility of oxygen when the temperature gets much hotter in the summer? Why have fountains been installed in some newly made lakes and ponds at apartment complexes, subdivisions, and golf courses?