

9.9: Gay-Lussac's Law

A third gas law may be derived as a corollary to [Boyle's](#) and [Charles's](#) laws. Suppose we double the thermodynamic temperature of a sample of gas. According to Charles's law, the volume should double. Now, how much pressure would be required at the higher temperature to return the gas to its original volume? According to Boyle's law, we would have to double the pressure to halve the volume. Thus, if the volume of gas is to remain the same, doubling the temperature will require doubling the pressure. This law was first stated by the Frenchman Joseph Gay-Lussac (1778 to 1850). According to **Gay-Lussac's law**, *for a given amount of gas held at constant volume, the pressure is proportional to the absolute temperature*. Mathematically,

$$P \propto T \text{ or } P = k_G T \text{ or } \frac{P}{T} = k_G$$

where k_G is the appropriate proportionality constant.

Gay-Lussac's law tells us that it may be dangerous to heat a gas in a closed container. The increased pressure might cause the container to explode, as you can see in the video below. The video shows very, very cold nitrogen gas in a bottle being warmed by the air. Since the bottle's volume is relatively constant, as the temperature of the nitrogen gas (formed when the liquid nitrogen boils) increases, so does the pressure inside the bottle until, finally, BOOM!



✓ Example 9.9.1: Temperature

A container is designed to hold a pressure of 2.5 atm. The volume of the container is 20.0 cm³, and it is filled with air at room temperature (20°C) and normal atmospheric pressure. Would it be safe to throw the container into a fire where temperatures of 600°C would be reached?

Solution

Using the common-sense method, we realize that the pressure will increase at the higher temperature, and so:

$$P_2 = 1.0 \text{ atm} \times \frac{(273.15 + 600) \text{ K}}{(273.15 + 20) \text{ K}} = 3.0 \text{ atm}$$

This would exceed the safe strength of the container. Note that the volume of the container was not needed to solve the problem.

This concept works in reverse, as well. For instance, if we subject a gas to lower temperatures than their initial state, the external atmosphere can actually force the container to shrink. The following video demonstrates how a sample of hot gas, when cooled will collapse a container. A syringe barrel is filled with hot steam (vaporized water) and a plunger placed to cap off the end. The syringe is then placed in a beaker of ice water to cool the internal gas. When the temperature of the water vapor decreases, the pressure exerted by the vapor decreases as well. This leads to a difference in pressure between the vapor inside the barrel and the atmosphere. Atmospheric pressure then pushes the plunger into the barrel.

This page titled [9.9: Gay-Lussac's Law](#) is shared under a [CC BY-NC-SA 4.0](#) license and was authored, remixed, and/or curated by [Ed Vitz, John W. Moore, Justin Shorb, Xavier Prat-Resina, Tim Wendorff, & Adam Hahn](#).