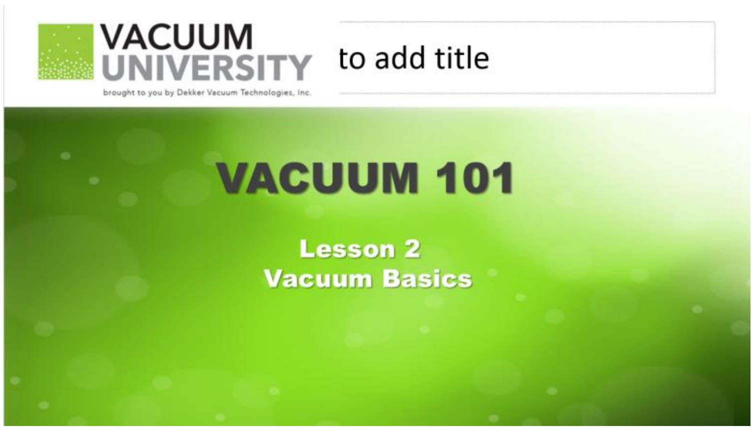
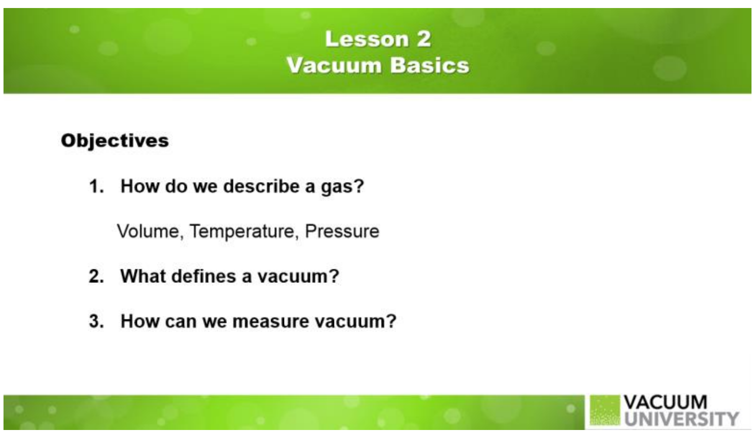
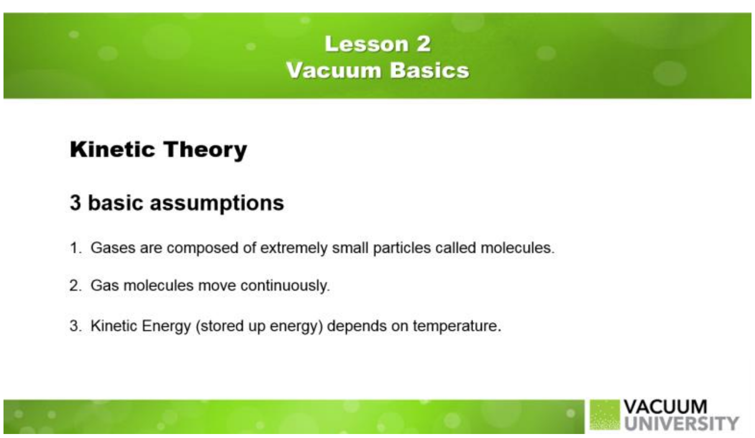
Vacuum University 101

Lesson 2: Vacuum Basics

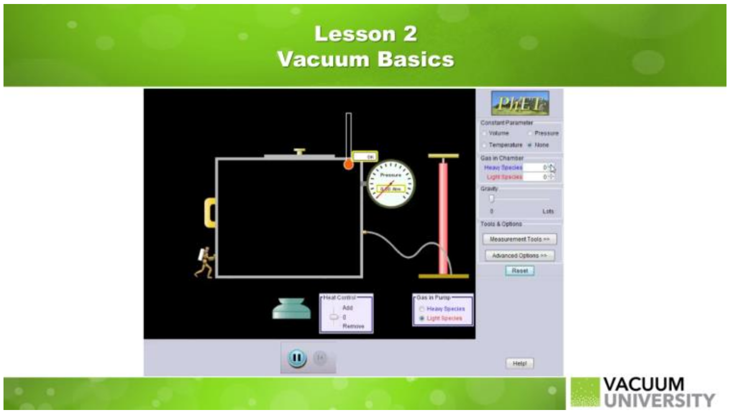
<table>
<tr><th>Image</th><th>Voice Narration</th></tr>
<tr><td>


</td><td>
Welcome to Vacuum 101, Lesson 2, Vacuum Basics. In this module, we will discuss gases, including air and their significance to understanding vacuum.
</td></tr>
<tr><td>


</td><td>
Understanding vacuum requires a familiarity of the properties of gases.

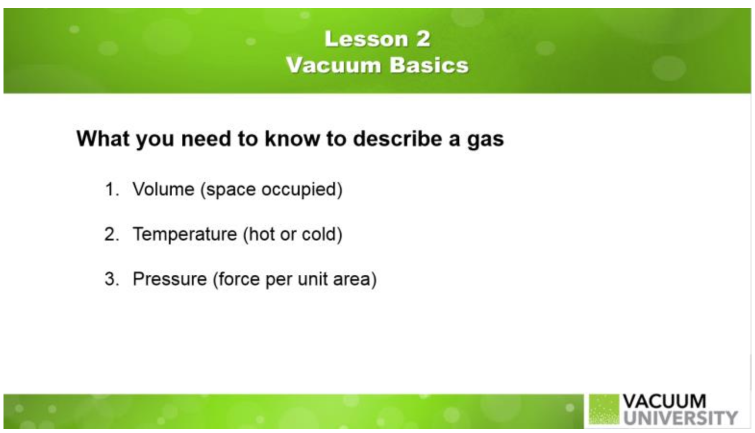
The objectives for this module are to describe a gas in terms of volume (the space it occupies), temperature (in degrees), and pressure (using various units of measurement). Then we will explore what defines a “vacuum” or very low pressure, and how we can measure vacuum.
</td></tr>
<tr><td>


</td><td>
You might recall from High School Physics the term ‘Kinetic Theory. Kinetic Theory is credited with the creation of Vacuum Science. Kinetic Theory states that gas pressure is a result of molecules impacting a surface; that the ‘pressure’ is equal to the rate of change of the mass of the molecules and the speed of the molecules. Simply put, Kinetic Theory comes from three basic assumptions:
1) Gases are composed of extremely small particles called molecules.
2) Gas molecules are in continuous motion as they move in straight lines. Upon colliding with the container wall, they behave as perfectly elastic spheres bounding all about. When they do, they exert “pressure” on the surface of the container.
3) The Kinetic Energy (stored up energy) of molecules is dependent on the temperature of the gas. The higher the temperature, the faster the molecules move.
</td></tr>
</table>



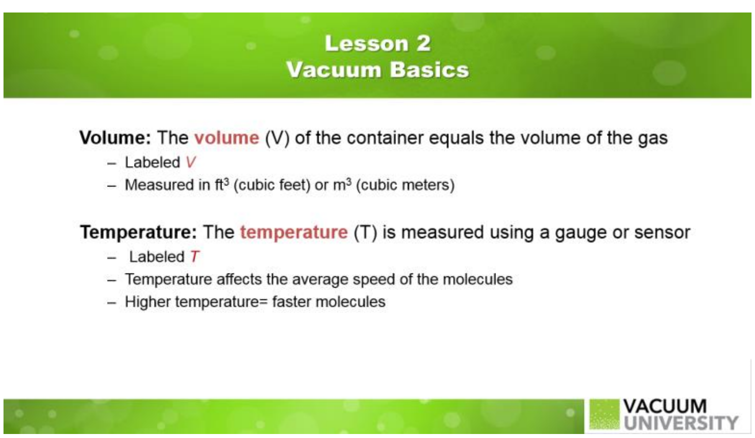
Let’s imagine our gas as we described as a collection of molecules. We could assume these molecules are like tiny tennis balls, flying around and bouncing off each other, and any surfaces that they come into contact with.

[Start demonstration]
Here is a simulation that will demonstrate gas properties. We introduce molecules into this container. Observe the interaction of the molecules with their surroundings. Just like molecules in the air or atmosphere, they bounce off surfaces and each other. Although the simulation shows molecules contained in a vessel, molecules (such as Nitrogen and Oxygen) are in fact around us all the time.



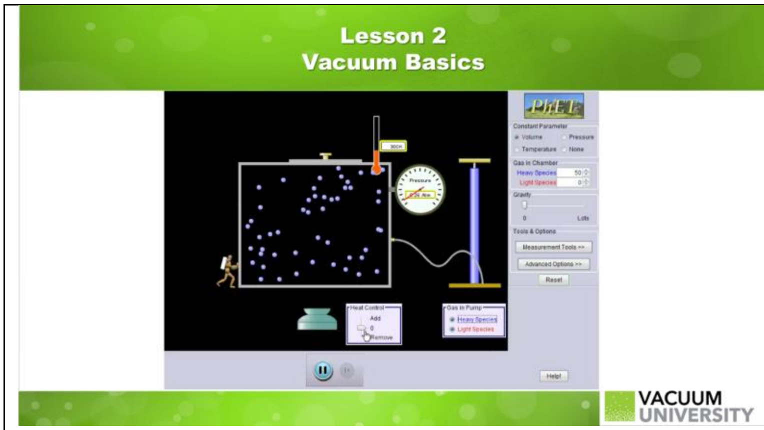
In order to describe a gas (other than what it is, like Nitrogen or Oxygen), you will need to know several properties of that gas.

You will need to know how much “space or volume” is occupied by the gas, the temperature of the gas, and the pressure exerted.
There are other properties of gases that are required for calculations which will be presented in later lessons.



Let's look at volume and temperature first. The volume of the gas (designated as “V”) equals the volume of the container and is typically measured in cubic feet or cubic meters.

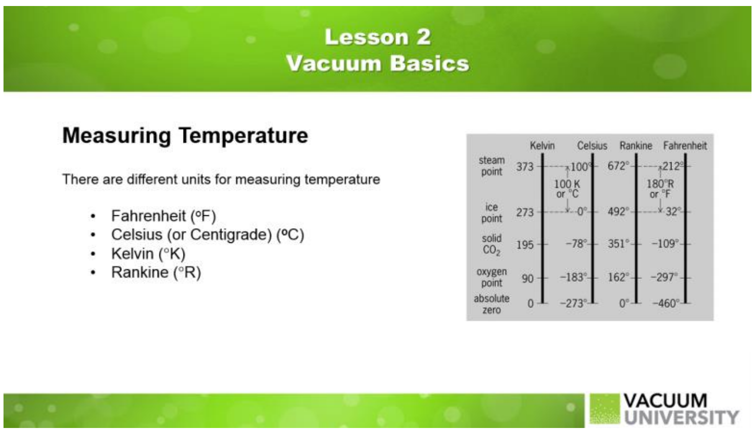
Temperature (designated as “T”) is another important property to consider. Measured using a gauge or sensor, temperature affects the Kinetic Energy of the molecules.



Let’s demonstrate. [Start Demonstration]

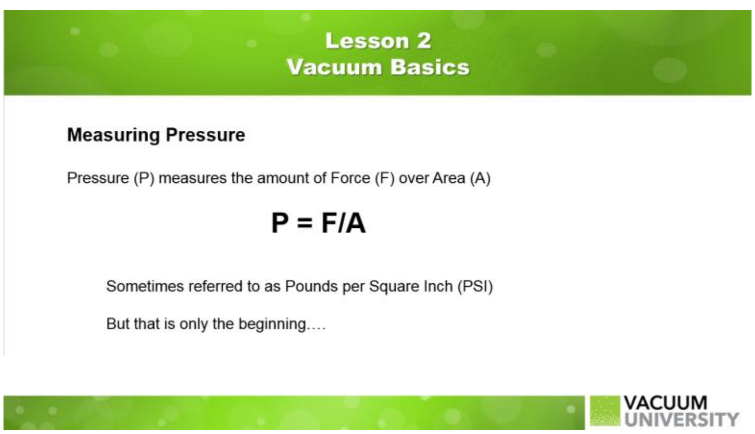
There are already molecules present in this container. Let’s raise the temperature What do you think will happen?

Notice, as the temperature goes up, the molecules increase in speed. When we reduce the temperature in the container, what happens to the molecules? That's right, they slow down. So temperature has a dramatic effect on gases.

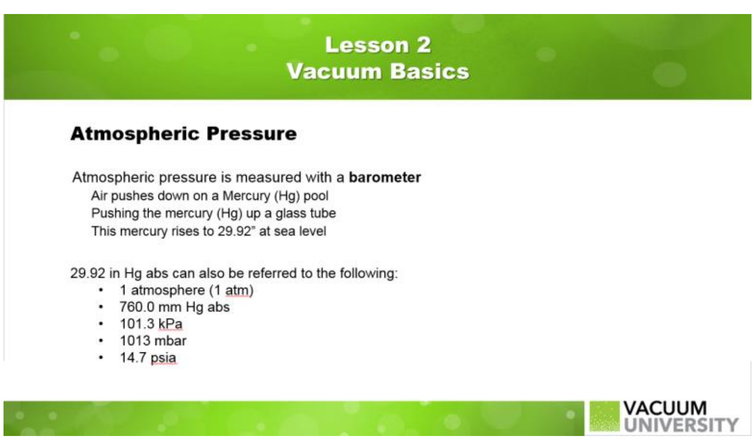


How do you measure temperature? Typically, you've heard of Fahrenheit or Celsius. But there are also scales called Kelvin and Rankine that are important, especially as we get into measuring temperature in a vacuum.

Kelvin and Rankine are related to Celsius and Fahrenheit, respectively. It's just that they are adjusted relative to absolute 0 degrees. That’s their starting point. The scales are set up so that 0^o Kelvin and 0^o Rankine mean that molecules have no velocity and therefore no kinetic energy. Sometimes this point is referred to as absolute 0^o.

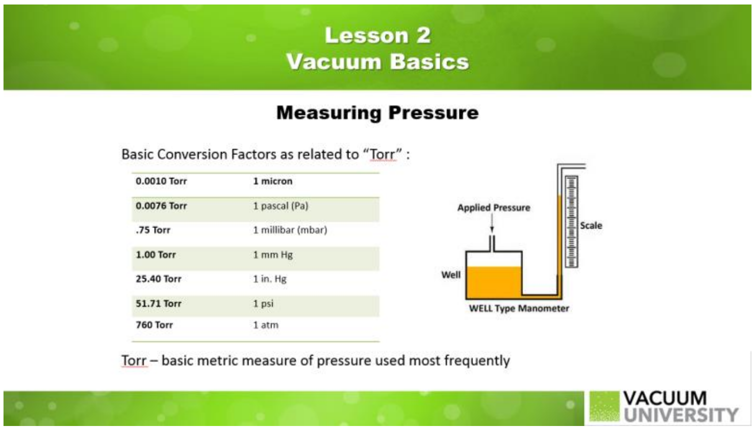


Pressure is another important element in understanding gases. When molecules collide with a surface, they exert force. Pressure (designated as “P”) basically measures the amount of force per square inch or square foot (or meter) of area. Based on this, we can establish a formula where pressure equals force divided by area. Therefore, higher pressure means equal molecules are hitting harder and more often. One of the more common terms to describe pressure is PSI or pounds per square inch which is a unit of measurement used in the United States. But you will see that this is just the tip of the iceberg when it comes to measuring pressure.

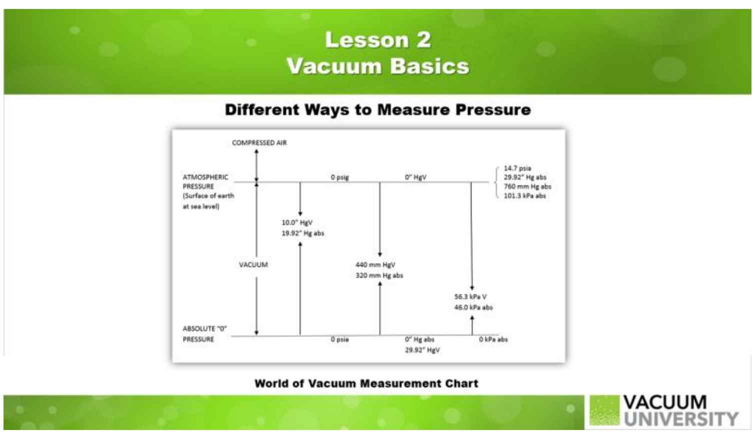


Recall in our previous lesson we discussed a barometer as a tool used to measure pressure. A simple barometer consists of a long glass tube (closed at one end, open at the other) filled with mercury and turned upside down in a container of mercury. The barometer works by balancing the mercury in the glass tube against the outside air pressure. As the air pressure increases (the air gets heavier), it will push more of the mercury up into the tube. As the pressure decreases, more mercury drains from it. So, the level of mercury in the tube provides the measure of air pressure.

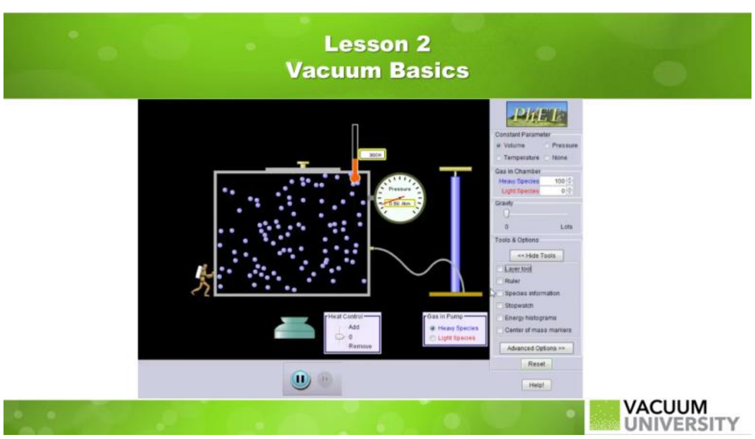
To simplify, 29.92 inches of mercury can also be referred to as 1 atmosphere. It is also measured as 760 millimeters of Mercury when using metric or 760 Torr (remember Torricelli and the use of Torr?). Other terms include 101.3 kPa, 1013 mbar or as we mentioned before, 14.7 psi.



Since Torr is one of the most common measurements of pressure or vacuum, a conversion chart like the one you see here is sometimes used to keep things straight.

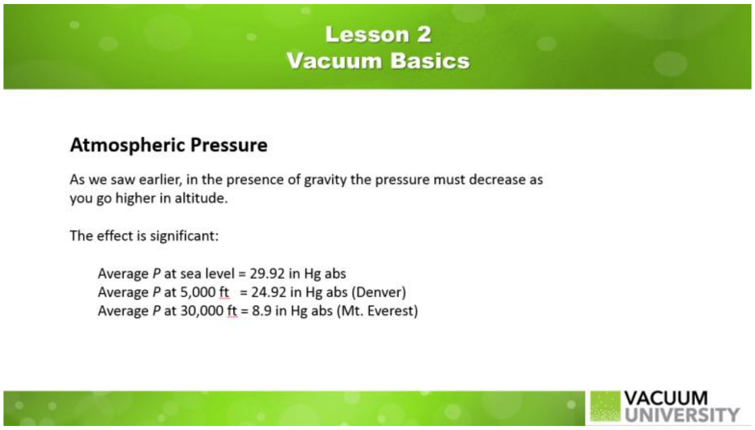


One more important consideration when measuring pressure is to determine if you are measuring absolute or gauge pressure. Absolute pressure is when referencing against a perfect vacuum, so it is equal to atmospheric pressure minus gauge pressure. Gauge pressure is equal to atmospheric pressure minus absolute pressure. Negative signs are usually omitted. To distinguish a negative pressure, the value may be appended with the word "vacuum" or the gauge may be labeled a "vacuum gauge." An absolute pressure gauge shows the actual pressure in a closed vessel independent from the atmospheric pressure.



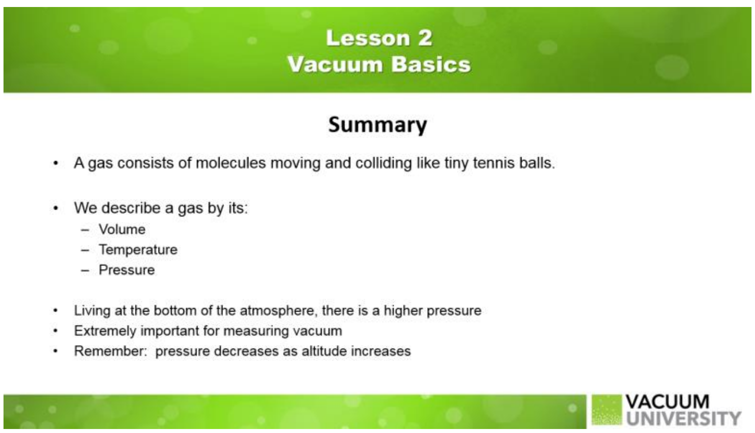
This demonstration shows the effect of gravity which will mimic our atmospheric conditions which are essential for understanding how to measure pressure or vacuum. [Start Simulation]

Let's provide some gravity inside of this container. As the gravity increases, notice what happens to the molecules. They do indeed reside mostly at the bottom, but there are still several bouncing towards the top. In fact, when measuring the pressure at the top of the container, you'll notice it's very low. But at the bottom of the container, where it's more dense, you'll notice the pressure goes up.



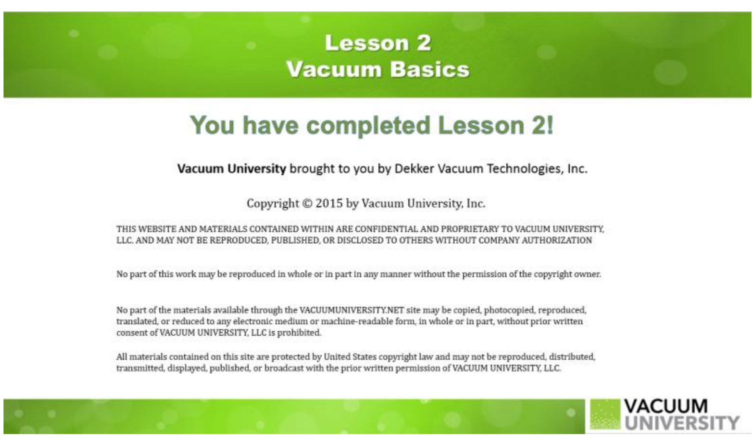
So, we noticed that the pressure decreases as you go higher in altitude. This effect is very significant when measuring pressure. Because the average atmospheric pressure at sea level, remember, is 29.92 inches mercury absolute.

But the average pressure at atmosphere, when you're going up to 5,000 feet, let's say, in Denver, it is quite a bit lower at 24.92 inches mercury absolute. And at Mount Everest, 30,000 feet high, it's at 8.9 inches mercury absolute.



In summary, a gas consists of molecules moving around and colliding like tiny tennis balls. We've described a gas by: Volume, Temperature, and most importantly, Pressure. Living at the bottom of the atmosphere, there's a higher pressure.

It might often be unnoticeable. You might not think about it, but it's extremely important to know this to better understand the conditions for measuring vacuum. Recall –Pressure decreases as the altitude increases.



This concludes Vacuum 101, Lesson 2: Vacuum Basics. Brought to you courtesy of Dekker Vacuum Technologies, Inc.