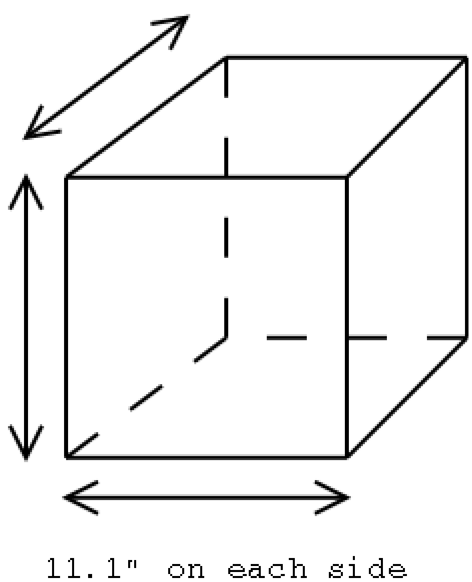
Introduction

We've discussed before how the properties of different gases are so similar that they obey the same laws no matter what the chemical identity of the gas itself is. What we **haven't** done yet is to relate the properties of gases (pressure, volume, and temperature), to the actual amount of gas present.

You might suspect that the amount of a gas present could be determined if you knew the volume of the gas and its pressure and temperature. You'd be right. Experimentally, the same number of moles of gas take up the same space, no matter what the gas actually **is**. This relationship is called Avogadro's Law, and it sets the stage for a useful relationship between the moles of a gas and the gas's properties.

Avogadro's law: The molar volume

To get a feel for the way gases behave, you need a sense of perspective. How much space a gas actually takes up in conditions that you're familiar with? Experimentally, at **0 ºC** and **1 atm pressure** (in other words, 32 ºF and the pressure at sea level) a **mole of gas** takes up a volume of **22.4 L**. That's a little over eleven two-liter soft drink bottles.

 	*Standard Temperature and Pressure:* • **1 atm pressure** • **0 ºC (273K)** *At these conditions, a gas occupies* • **22.4 L per mole of gas molecules** • The volume of a mole of gas (the **molar volume**) changes as we change the temperature or pressure of the gas! • The volume does **not** change based on the **chemical identity** of the gas!
Illustration 1 - The molar volume in English units	

These conditions are known as **standard temperature and pressure**, or **STP** for short. The volume at STP is a common way for scientists and engineers to describe an amount of gas..

The Ideal Gas Law

From Avogadro's law and our observations on the molar volume, we conclude that there is probably an equation that relates the moles of gas molecules present to the other measurable properties (pressure, volume, and temperature) of a gas. This equation was discovered by experiment just like the molar volume and the other empirical gas laws:

$$PV = nRT$$

... where:

P = pressure (in atm)	R = 0.08206 L·atm / mol·K (a constant)
V = volume (in L)	T = temperature (in K)
n = the number of moles	

This equation is called the **ideal gas law**, and it has been found to hold true for gases at normal conditions. (We will discuss gases at extreme conditions later.)

Using the ideal gas law, we can tell how many moles of gas molecules occupy a given space. Or we can go the other way around and find out how much space some number of moles of gas would require at some set of conditions. Many reactions produce gases as products, and if too much gas is released into a confined space, an **explosion** can occur! Some common explosives are solids that decompose or react rapidly to form many gas molecules, and we will look at an explosive, **ammonium nitrate**, in detail in a future note pack.

Let's look at a quick example of the ideal gas law.

How big is a pound of nitrogen gas, N_2, at room conditions? (25 ^oC and 1.00 atm)

Strategy: If we're given information about the mass or moles of a gas (or are asked about either), we will need to use the ideal gas law. We've been given the mass of gas present, but it's in an undesirable unit - pounds. As scientists, we prefer grams. Since one pound is equal to exactly 454 grams, we will find the volume of 454 g of N_2 gas.

First, find moles of gas via the formula weight:

$$454\ \text{g}\ N_2 \times \frac{1\ \text{mol}\ N_2}{28.02\ \text{g}\ N_2} = 16.20\ \text{mol}\ N_2$$

Now, we can use the ideal gas law.

$$PV=nRT$$

P = 1.00 atm *(specified in problem)*
V = ??? *(what we've been asked)*
n = 16.20 mol *(found above)*
R = 0.08206 L·atm / mol·K *(a constant)*
T = 298 K *(convert 25 °C to K)*

Plug in and solve. The simplest way to do this is to solve PV=nRT for V and **then** plug in!

$$V = \frac{n \times R \times T}{P}$$

$$V = \frac{(16.20\ \text{mol}) \times (0.08206\ \text{L} \cdot \text{atm} / \text{mol} \cdot \text{K}) \times (298\ \text{K})}{(1.00\ \text{atm})} = \mathbf{396\ L}$$

V = **396 L occupied by a pound of N_2**.

So, one pound of nitrogen gas takes up a rather large volume at room conditions. This makes sense, since we know that gases aren't very dense.

Most of the units cancel out to leave you with a volume unit. It's particularly important when working with the ideal gas law to make sure you have your numbers in the right units before plugging in! Your gas volume, pressure, and temperature should have the same units as the value for the ideal gas constant, R.

Next, we will discuss a theory that accounts for the ideal gas law. We will then use the ideal gas law to deal with stoichiometry problems involving gases.

<u>Summary</u>

A mole of gas has been observed to take up 22.4 liters of volume at a certain set of conditions (STP). This volume was effectively independent of the chemical nature of the gas. Oxygen, nitrogen, hydrogen, and other gases all had a molar volume of 22.4 L at STP. This observation led to the discovery of the ideal gas law, relating gas properties - pressure, volume, and temperature - to the moles of gas molecules present. For us, that means we can now figure out the volume of gas produced in a chemical reaction. A real-world application of gas volumes and reactions is an observation of the nature of explosives - chemical reactions that produce a large volume of gas quickly from a small volume of solid starting materials.