Chapter 14: Let Loose the Gases!

Gases can influence all aspects of our lives, from the air we breathe to the air we wish we didn't have to breathe. This store is probably someplace you don't want to buy a sandwich.

Photo by David Jordan, used with permission: http://commons.wikimedia.org/wiki/File:Mood_fart.png

Chapter 18: Let Loose the Gases!

Gases are really, really important. The tire in your car wouldn't stay inflated unless it was full of gas. The balloons at your birthday party wouldn't stay inflated if it weren't for the helium they put in them.[1] The ping pong balls you use to keep you entertained are full of air, which probably has something to do with their bounciness, though I'm not really sure. Gases are all over the place![2]

Bringing Science Home

To test the importance of gases, perform the following mini-lab:

1) *Put strips of duct tape over your friend's nose and mouth.*
2) *Observe.*

Post-lab questions:
What happened to your friend? Do you think this had something to do with the lack of air? Have you given any thought to what you're going to tell the police?

Figure 14.1 credit:
http://commons.wikimedia.org/wiki/File:Cling_duct_tape.jpg

Section 14.1: The Combined Gas Law

As we mentioned in Chapter 13, the properties of gases can be explained by the kinetic molecular theory. If you're not sure what this theory says, go back and reread that section.

OK… you're back. Now that you know why gases have their particular properties, let's talk about some of the mathematical expressions that can describe how they behave.[3]

Boyle's Law

Boyle's law says that when you squish a gas, its pressure goes up. Usually expressed as $P_1V_1 = P_2V_2$, Boyle's law just says that the pressure and volume of a gas are inversely proportional to each other (i.e. as one goes up, the other goes down).

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[1] Or, if your family is cheap, they just blow it up with their breath. Cheapskates.
[2] If you haven't realized that the lab below is a joke, be warned that it is and that you shouldn't kill your friend. Or if you do, use a method that can't be traced back to this book. I'm not going back to prison.
[3] In each equation, you'll see things like P_1 and P_2 (and corresponding terms for V and T). This refers to the fact that we're measuring how these variables are altered after a change: P_1 is the pressure of the gas before it undergoes a change, while P_2 is its pressure after it undergoes the change.

Charles's Law

Charles's law states that when you increase the temperature of a gas[4], it gets bigger. This is usually attributed to the fact that the particles in a hot gas fly around at higher speed, causing them to hit the side of their container harder and push it outward. If you want to see the math version of this equation, check this out:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Important: Temperature is measured in Kelvin and not ^{0}C. K = ^{0}C + 273

Gay-Lussac's Law

If you're giggling at this guy's name, grow up. It's not that funny, particularly when you realize that the name "Gay-Lussac" means "Destroyer of Men's Souls" in French. Though it has never been verified, I hear that Gay-Lussac totally killed like fifty guys with his bare hands because they made fun of his clothes.

Figure 14.2: *Though, in all fairness, his clothing looked a little stupid, even for the time in which he lived.*

http://en.wikipedia.org/wiki/File:Gaylussac_2.jpg

Anyway, Gay-Lussac said that if you increase the temperature of a gas that's stuck in a sealed container, the pressure of the gas will increase. The idea here is that if you increase the speed at which the particles travel, they'll hit the side of the container at a higher speed, causing the pressure inside to go up. For those of you who like equations, Gay-Lussac's law is numerically expressed as:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Important: Temperature is measured in Kelvin and not ^{0}C. K = ^{0}C + 273

A Handy Term You Should Probably Know

Figure 14.3: *The term STP stands for "standard temperature and pressure" and refers to a temperature of 273 K (0^0 C) and a pressure of 1 atm (101.325 kPa). It also makes reference to the "St. Paul Sandwich" shown here. Yum!*

http://commons.wikimedia.org/wiki/File:STP_Sandwich_005.jpg

[4] In all gas law calculations, always measure temperature in Kelvin rather than ^{0}C. In case you forgot, K = ^{0}C + 273.

The Combined Gas Law (The Important One)

Let's say that, like most of the human beings in the world, you don't want to memorize all three of the laws above. In that case, you need only memorize one gas law that puts all three of the other laws together into a big convenient blob:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Important: Temperature is measured in Kelvin and not ^0C. $K = {^0C} + 273$

You only really need to memorize this law instead of the other three because it says the same thing. For example, if a problem doesn't mention the temperature of a gas, just leave it off and you're left with $P_1V_1 = P_2V_2$, which is Boyle's law. If a problem doesn't mention the pressure of a gas, you're left with $V_1/T_1 = V_2/T_2$, which is Charles's law. If a problem doesn't mention the volume of a gas, you're left with $P_1/T_1 = P_2/T_2$, which is Gay-Lussac's law.

Focus on Chemistry Stuff

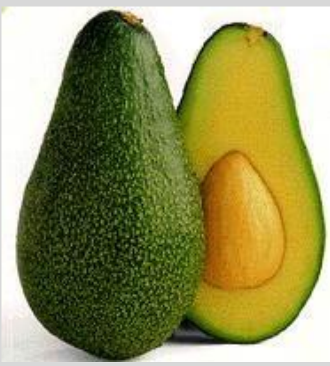

Figure 14.3: *I think it's funny to think of it as Avocado's law.*

*The kinetic molecular theory of gases describes the properties which are common to all gases. Because it states that all gases have exactly the same properties, all gases are said to have the same molar volume (i.e. volume per mole of the compound) – this is called **Avogadro's law**. Thus, 1 mole of methane will be expected to have the same volume as 1 mole of oxygen. **Ideal gases** are gases which actually behave like this (i.e. none of them), while **real gases** are the gases from the real world that don't exactly follow the KMT. Still, real gases behave closely enough to ideal gases that these gas laws are still pretty good approximations of how gases behave.*

http://commons.wikimedia.org/wiki/File:Avocado.jpeg

Section 14.2: Ideal Gases

In this section, I'm going to discuss the "ideal gas law", which we usually use as a guide to explaining how the volume, temperature, and pressure of a gas are related to the number of moles present. However, before I go into the details of the ideal gas law, it's probably best if I explain what an ideal gas is.

What's an Ideal Gas?

When I talked about the kinetic molecular theory in the last chapter, I mentioned that it was a series of approximations that was used to explain the behavior of all gases. However, it's worth mentioning that there's actually no such thing as an ideal gas, because these approximations are just that – approximations of how gases actually behave. For example, gases do experience intermolecular forces, but they're very weak because the gas particles are moving so quickly and usually aren't very close to each other. Though we treat gas molecules as being infinitely small, they actually have a very small volume so they deviate from the predictions of the KMT. Gas molecules also don't undergo elastic collisions, because a small amount of energy is lost when they collide with other things. Though all of these effects are small, they do cause gases to behave slightly differently than the KMT would predict.[5]

That said, most gases come very close to behaving like the KMT says, and the theoretical gas that conforms to the postulates of the KMT is referred to as an **ideal gas**. Again, there are no ideal gases in the real world – only **real gases** that behave in a similar way.[6]

The Ideal Gas Law

The ideal gas law is an expression that describes the relationship between the pressure, volume, temperature, and number of moles of a gas. To put this in numerical terms, it states that:

$$PV = nRT$$

Where:

- P = pressure (in atmospheres or kilopascals)
- V = volume (in liters)
- n = number of moles of gas
- R = the ideal gas constant, which is either 0.08206 L·atm/mol·K or 8.314 L·kPa/mol·K[7]
- T = temperature (in Kelvin)

For example, let's say that we have 4.00 moles of carbon dioxide at a temperature of 25^0 C at a pressure of 1.50 atm. Given this information, what is the gases volume?

To answer this, just plug the numbers into the ideal gas law. In this case, the pressure is 1.50 atm, the volume is unknown (X), the number of moles is 4.00 mol, the value of R used is 0.08206 L·atm/mol·K, and the temperature is 298 K (remember, we need to convert all temperatures to Kelvin when performing gas law problems).[8] Plugging these values into the ideal gas law, we find that:

$$(1.50 \text{ atm})X = (4.00 \text{ mol})(0.08206 \text{ L})(298 \text{ K})$$
$$X = 65.2 \text{ liters}$$

About the only thing that can make this problem more difficult is if the quantity of gas is given to you in grams. If this occurs, just do a simple mole calculation to convert it into moles.

[5] Additionally, under conditions of low pressure and high temperature, real gases behave more closely to ideal gases. This is because these conditions cause particles to be far away from each other and to move at high speeds, both of which minimize the attractive effects of intermolecular forces.
[6] The similarity to which a real gas behaves as an ideal gas depends on a number of factors. Gases that are large and have strong intermolecular forces generally behave in a less ideal manner than small molecules with weak intermolecular forces.
[7] The exact value of R depends on the units given for pressure. If pressure is given in atm, the 0.08206 L·atm/mol·K is used. If the pressure is given in kPa, the 8.314 L·kPa/mol·K is used.
[8] In case you forgot, K = ^{0}C + 273.

Section 14.3: Gas Stoichiometry

You know that chart I gave you to learn about stoichiometry back in Chapter 12? Well, it turns out that, with a few small modifications, you can use it to do stoichiometry with gases. Let's take a look at the new version:

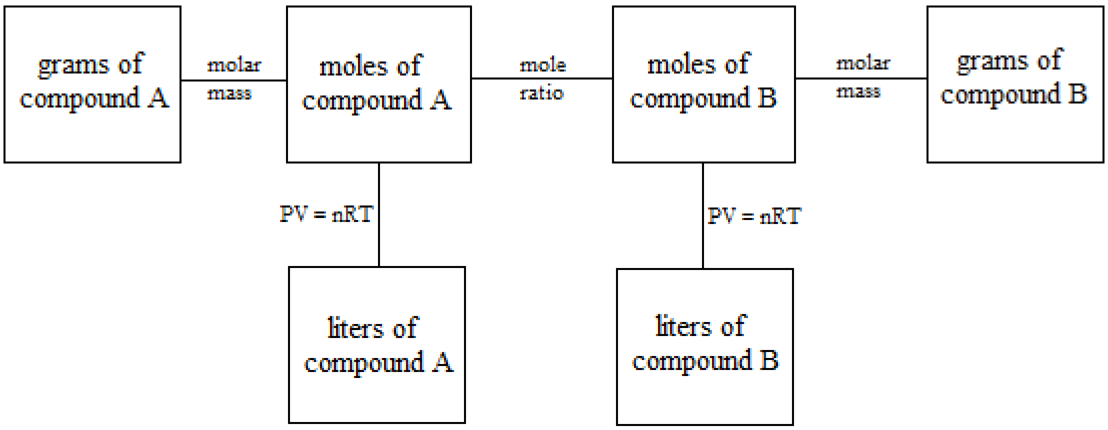

Figure 14.4: *The handy chart that we use to do stoichiometry, with information for using gases added.*

This table is used in exactly the same way as the other stoichiometry table, except that if you're given liters of a gas, or told to convert to liters of a gas, you have to use PV=nRT to convert between moles and liters.[9]

A Gas-Related Illusion:

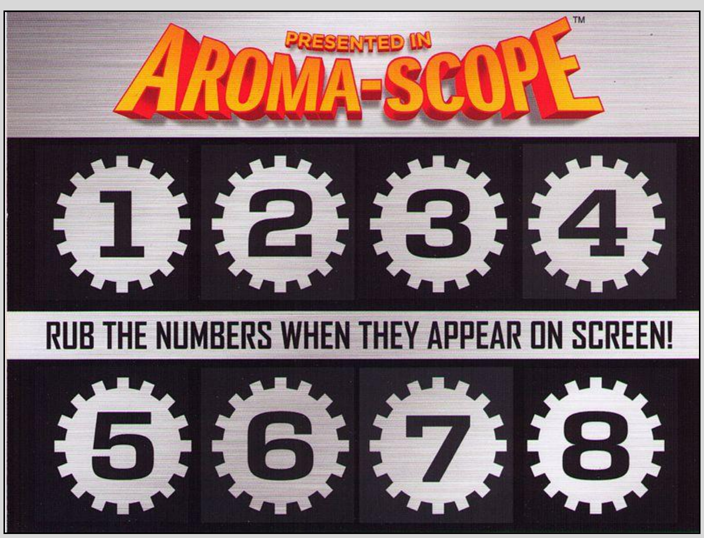

Figure 14.5: *The smells that reach your nose are really airborne chemicals released from the thing that gives off the odor. This illusion, the "Aroma-scope" is able to trick your brain into believing it smells various scents. Give it a shot!*

http://commons.wikimedia.org/wiki/File:Aroma-scope.jpg

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[9] Handy tip: At standard temperature and pressure (0^0 C and 1.00 atm), there are 22.4 liters of gas in 1 mole. This should help make these calculations easier if you're working at STP.