Podcast Script – Basic Gas Laws.

Hello,

This podcast comes to you today from Dhahran, Saudi Arabia, sponsored by Scramling Science. Today’s topic is Basic Gas Laws.

Around the early 1600’s, scientists started experimenting with gases. Not a whole lot was known about gases before this time for a few reasons. The gases were mostly invisible. Another problem was that before this, it wasn’t easy to contain the gases in a vessel. About the same time, glass working, esp. in Italy became quite good and with ground glass beakers and stoppers, gases could be more easily quantified and measured.

This means that pressure and its effects on the conditions of a gas can more studied more easily.

You will be introduced here to several formulas. They are comparing a change in a variable when another variable is being manipulated. It is important to keep in mind that the other variables not mentioned in the problem are being kept constant. For example, we will soon talk about Boyle’s Law which compares Pressure and Volume. We are assuming that the moles of gas and their temperature are being kept constant.

Sometimes in problems, the term standard pressure (SP) might be used instead of a numerical value. Remember that standard pressure is 1 atm., or 101**.**325 kPa or 760 mmHg. Standard Temperature (ST) is the same as 0 °C or 273 K. Sometimes these are combined to be called Standard Temperature and Pressure (STP).

Around 1650, an Irish alchemist named Robert Boyle, was working with gases with some of the new glassware available. He was experimenting with the pressure of a system of gases and determining the volume. Boyle believed in experimentation through a designed and logical method, which at the time wasn’t usually practiced. He is often considered the first true chemist. His noticed through his experiments, that pressure and volume were indirectly proportional and so P times V always equaled a set constant for that sample. His equation started as PV = k where k is a constant.

Since the value of k was always a constant for that sample, two sets of conditions of the gas can be compared and Boyle ended up with P1 V1 = P2 V2  Remember that the units must always match up.

About 100 years after Boyle, a French scientist named Jacques Charles determined that volume and temperature were directly proportional. Actually, it was known for quite some time that as temperature increased, so did the volume. However, Charles’ took the work of an English scientist named Lord Kelvin who was working on a new way of measuring temperature which didn’t depend on any particular physical property. The problem with the Celsius temperature scale was the following: work this example to see for yourself.

If the volume of a balloon at 10 °C was 1**.**4 Liters, what would be the volume at – 10 °C. Clearly the value of – 1**.**4 Liters does not make sense. The experiment was clear but the results needed some interpretation. Using the Kelvin scale and the same formula of V1 / T1 = V2 / T2 we get a result of 1**.**3 Liters which was verified by experimentation.

The question might arise that what if both pressure and temperature are changing – what will then happen to the volume. It is possible to take two of the variables and combine them algebraically. This was done and the Combined Gas Law was created. It is used for Pressure, Temperature, and Volume where two of the three variables are constant and the third is being solved for. The formula is P1 V1 / T1 = P2 V2 / T2

Amadeo Avogadro is known for his work with the mole but his initial experimentation dealt with gases. His work showed that the volume of a gas was directly proportional to the number of moles of the gas. The formula was n1 / V1 = n2 / V2

Gay-Lussac’s Law shows the relationship between Pressure and Temperature. We have discussed Pressure on occasion on a few of these Podcasts. One way to think about Pressure is to talk about it in terms of gas particles coming into contact with the interior walls of the container it is in. Consider what happens when matter is heated. Well, of course it gets hotter, and temperature is defined as the average Kinetic Energy of the molecules of a system. Since the molecules are moving faster, on average, they are going to hit the insides of the container more frequently. As a result the pressure goes up.

Gay-Lussac’s formula is P1 / T1 = P2 / T2

If you examine all of the formulas above, you notice that they all depend on changing conditions of the variables being measured. This is fine, if you can do that. Sometimes the variables aren’t changing, such as if the gases are in a rigid container like a steel cylinder. What was needed was a formula that could be used when the conditions were not changing.

Some resourceful person noticed that if there was a sample of gas and you took the Pressure multiplied by the Volume and then divided that by the moles multiplied by the Temperature in Kelvin, it was always equaling a constant. This idea was algebraically manipulated and the result was the Ideal Gas Law which is written as PV = nRT. R is the Gas Law Constant which is the ratio of PV / nT stated above. This now gave a method to calculate any property of a gas when the others were given.

The algebra is pretty straight-forward. Any variable can be the unknown that is being solved for, and you will be doing plenty of examples in class. A hurdle to look out for is that the units must always match up. Since pressure can easily be in either atmospheres, kPa, or torr, the value of R changes depending on the unit used for Pressure. Remember that the actual concept doesn’t change and to a point it seems counter-intuitive that a constant has a value that changes, but it really isn’t changing itself, it is just undergoing a unit conversion.

The values for R that you should know are: **.**08205 atm. L / n K Another one is 8**.**314 J / n K you notice that there doesn’t seem to be a pressure unit present here, but the Joule (symbol J) contains the kPa for pressure. The value of 8.314 is used more in Physics and certain times in Chemistry. When the unit for pressure is the torr, the value of R = 62**.**36 torr L / n K

In an earlier podcast we learned that there were standard conditions of a gas that are often used so that comparisons can be made. Recall that these are 1 atm. and 0° C and are often called STP, which stands for Standard Temperature and Pressure. Taking these values, we can obtain a new helpful value and that is the volume of one mole of gas. Substituting in values for PV = nRT gives us the following:

1 atm. x V = 1 mole x **.**08205 x 273 K. Be careful with units and notice that since we used a standard pressure of 1 atmosphere, we needed to use **.**08205 as the value of R to make the units match up. Temperature for all gas law formulas must be in Kelvin. Solving gives us 22**.**4 Liters and this is called the Molar Volume of a Gas. 1 mole of any gas occupies 22.4 Liters of volume at STP. This is an approximation, but a good one. It is used quite often.

It is also good to know that the Ideal Gas Law is not always exact. It is an excellent tool for when the gas is at low pressures and high temperatures, which is often the case. However, if the gas is at very high pressures or at low temperatures, the Ideal Gas Law breaks down quite a bit and we need to use the Real Gas Law which will be detailed in a later Podcast.

Dalton’s Law of Partial Pressure states the total pressure of a system is the sum of the pressures of the individual gases that make up the system. P total = P1 + P2 + P3 ….. Make sure that the units match up as this is the main problem when working these types of problems as the math is clearly quite simple. A good use of this type of problem is when you are doing water displacement to isolate a gas. Since gases are invisible, they can’t be seen. They are often obtained by doing a reaction in glassware with a tube attached that moves the gas to an inverted flask. The flask is filled with water and the gas displaces the water and fills it up. Since water will evaporate at any temperature, it must be factored in the calculations and this is done by Dalton’s Law of Partial Pressure. The pressure of the water vapor is subtracted out and what is left is the pressure of the dry gas.

I hope you enjoyed today’s Podcast and found it entertaining and educational, and helped you with your understanding of Basic Gas Laws. Remember that you can get more information on this topic from the class website or you can always send me a note on either FaceBook or via email. Refer back to this topic when needed, courtesy of Scramling Science.