Chem 1311Ch5Ep3 Transcript

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Hello, and welcome to the third episode of gases and gas laws.

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Previously, in gases and gas laws.

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We learned to derive both the ideal gas law and the combined gas law from Boyle’s, Charles’ and Avogadro’s gas laws.

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And we learned to apply both the ideal gas law and the combined gas law to solve word problems.

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In today's episode.

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We will derive the density and molar mass formulas for gases.

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And we will apply both formulas to problem solving.

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We will derive the density formula for gases from the.

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Ideal gas law,

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The definition of density, and the definition of molar mass.

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The ideal gas law shows the relationship between pressure, volume, number of moles and temperature for any gas.

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The definition of density.

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Is mass per unit volume.

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And the definition of molar mass.

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It's mass.

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Divided by the number of moles.

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We will start by solving the ideal gas law for the volume.

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And solving the definition of molar mass.

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For mass.

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We can then substitute.

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nRTP for the volume.

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In the definition of density.

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And we can substitute molar mass times the number of moles.

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For the mass.

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In the definition of density.

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This results in this combined formula.

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For density.

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Density will be equal to the pressure.

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Times the molar mass times the number of moles divided by the number of moles.

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Times the gas constant.

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Times the temperature.

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Notice that in this equation the number of moles cancels, so we don't need to worry about that.

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And now.

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We have the final version of the density for gases formula.

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Times molar mass divided by the gas constant divided by the temperature.

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This same formula.

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Can also be derived for the molar mass.

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And now we have both forms of this formula.

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For density calculations.

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Both of them are very useful, so you should know them both.

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An important distinction.

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The lowercase “m” is the mass in grams, whereas the italicized capital “M” is the molar mass in grams per mole.

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When dealing with gases, the density is measured as grams per liter.

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With other substances, solids and liquids, it is not.

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It's usually measured in grams per milliliter.

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But for gases it makes more sense to use grams per liter, since it gives you more reasonable values.

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Let's apply this new formula to a few examples.

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In this problem,

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The ask is the density of carbon dioxide.

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So, we will use the first form of the equation.

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They did not give us a molar mass, but we can figure it out from the formula.

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And we can convert the temperature to Kelvin.

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We make sure that all the excess units cancel.

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And our final density is 1.62 grams per liter.

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This next example is asking for the molar mass.

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Because of this, we will be using the second form of the density formula.

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We substitute-in the given values for the variables.

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We cancel all excess units.

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And we get a molar mass of 67.9 grams per mole.

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Ask for that second question.

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Let's see what we know about this compound.

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We know it's an oxygen chlorine compound.

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And we know its molar mass.

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We also know, using a little logic, that it can only have 1 chlorine.

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Because if it had two, that would put him.

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Over the molar mass.

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And therefore, it must have two oxygens.

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To achieve the required molar mass.

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It's not an exact fit.

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But it's pretty darn close.

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This next example we will have to split into.

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First, we will use the percent composition to get the empirical formula, just like we did back in Chapter 3.

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And then we will use the molar mass formula.

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To get the molar mass and molecular formulas.

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So next we divide the number of mole.

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By the smaller of the two values.

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And now we have an empirical formula.

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Silicone trifluoride, Let's see.

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It has an empirical formula mass of 85.1 grams per mole.

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So now to Part 2.

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Next, we are going to use the second form.

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Of the density formula to obtain the compound's molecular mass.

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We were not given the density.

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But we have our subtle ways to work around that.

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So we changed the formula just a smidge.

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And now we substitute the given values for the variables.

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Next, we can cancel all excess units.

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And we get 168.6 grams per mole.

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By comparing the formula mass to the molecular mass, we can conclude that the molecule is equal to the empirical Formula times 2.

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The last one of the gas laws we will be reviewing is Dalton's law.

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Dalton's law of partial pressures states.

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That in a mixture of gases.

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That are sharing a container.

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The total pressure.

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Is the sum of their individual pressures.

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This pressure of course will be a function of the gas mole fraction.

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Let's explain this a little bit.

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Here is a mathematical explanation.

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We can apply the ideal gas law to each of the individual components.

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In a mixture of gas A and gas B.

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We can also apply it to.

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Both of them together.

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For the total pressure.

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The pressure divided by the number of moles will be equal to RT / V.

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Regardless of whether you're looking at an individual gas or both gases together.

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Here we have 3 cases.

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Gas A, gas B, and the mixture.

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That also means that the pressure to number mole ratios in all three instances will be equal to each other.

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If we rearrange the equations slightly, we can see that the pressure due to gas A is equal to the total pressure.

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Times the mole fraction of A.

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This relationship is true for both component gases.

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A and B

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We have established that the total pressure is the sum of the pressures due to each individual component.

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That was our starting premise.

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And we have defined the mole fraction of each component gas as the ratio of the number of moles of that gas to the total number of moles.

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So it follows that the pressure due to gas A.

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Will be equal to the total pressure.

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Times the mole fraction of A.

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This, of course, is also true for gas B.

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And this will apply to any gas mixture.

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Here's an example problem.

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We will use Dalton's law and apply it to all three components.

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The total number of moles will be the sum of the number of moles of each gas, and it turns out to be 7.35 moles.

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The mole fraction for neon will be 4.46 mole divided by 7.35 moles or.

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0.607

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The same formula can be applied to argon.

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And to xenon.

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Once we know all three more fractions.

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We can now calculate each individual partial pressure by multiplying the total pressure.

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By the mole fraction.

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The sum of all three partial pressures should be equal to the total pressure of two atmospheres.

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This next example involves a mixture of oxygen gas and water vapor.

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The ask is the mass of oxygen.

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The first step is to use Dalton's law of partial pressures to determine the pressure of oxygen.

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Using this pressure.

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We can solve the ideal gas law.

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For the number of moles.

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We can then substitute the given values.

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And we can cancel the excess units.

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To obtain a final answer.

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Of 5.11×10^{-3} moles of oxygen.

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Which can then be converted to grams of oxygen.

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Doing a simple unit conversion.

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Using the molar mass value from the periodic table.

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This results in.

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0.163 grams of oxygen.

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The kinetic molecular theory of gases.

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Is the best explanation for all the unique properties of gases and the gas laws?

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It states that gas molecules occupy a negligible amount of space with vast distances between them.

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These molecules are in constant motion and frequently collide with each other.

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They also collide against the walls of the container, and these collisions are perfectly elastic.

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It also states that the molecules are neither, attracted, or repelled by each other.

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And the average kinetic energy of these molecules is proportional to the Kelvin temperature of the gas.

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At any given time,

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Lighter molecules will move faster than larger molecules.

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However, their kinetic energy will be the same.

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The high compressibility of gases is explained by the existence of vast amounts of empty space between the molecules.

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Boyle's law

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It's explained by the pressure being proportional to the collision rate.

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With a wall.

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This collision density in turn is proportional to the molecules’ population density, which in turn is proportional to the reciprocal of the of the container's volume.

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Charles law

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In turn, is explained because the collision rate is proportional to the kinetic energy of the molecules.

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Which in turn is also proportional to the temperature.

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Avogadro's law.

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Is explained because the collision rate is proportional to the number density.

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Which in turn is proportional to the number of moles.

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And Dalton's law of partial pressures.

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Is explained by the lack of attraction or repulsion between the molecules.

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And that's all there is.

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There isn't any more.