Molecules in Motion

Prof. Amita Pathak Mahanty
Department of Chemistry
Indian Institute of Technology, Kharagpur

Lecture – 04

Kinetic theory of gases (Contd.)

Welcome, to this course in Molecules in Motion. Today, we are going to be having 4th lecture that is Kinetic theory of gases. We will briefly go through what whatever we discussed in the previous class, so that we have a continuity.

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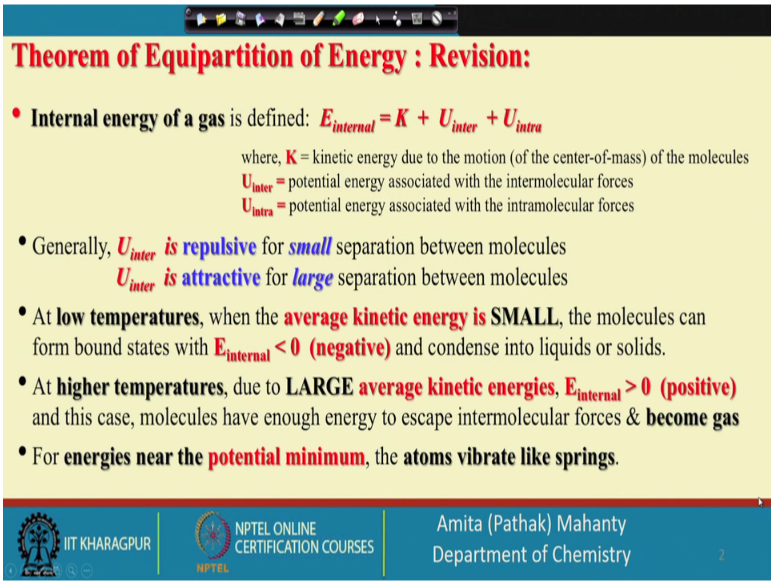

Theorem of Equipartition of Energy : Revision:

- **Internal energy of a gas** is defined: $E_{internal} = K + U_{inter} + U_{intra}$

  where, K = kinetic energy due to the motion (of the center-of-mass) of the molecules
  U_{inter} = potential energy associated with the intermolecular forces
  U_{intra} = potential energy associated with the intramolecular forces
- Generally, U_{inter} is repulsive for *small* separation between molecules
  U_{inter} is attractive for *large* separation between molecules
- At **low temperatures**, when the **average kinetic energy is SMALL**, the molecules can form bound states with $E_{internal} < 0$ **(negative)** and condense into liquids or solids.
- At **higher temperatures**, due to **LARGE average kinetic energies**, $E_{internal} > 0$ **(positive)** and this case, molecules have enough energy to escape intermolecular forces & **become gas**
- For **energies near the potential minimum**, the **atoms vibrate like springs**.

What we were talking about in the last class was the theorem of equipartition of energy. In this what we had talked about how the internal energy is distributed. Internal energy is distributed in the three forms; one is the kinetic energy of the motion of the molecule where we talk about the motion of the center of the mass and the internal potential energy because of the intermolecular forces or the potential energy associated with the intra molecular forces.

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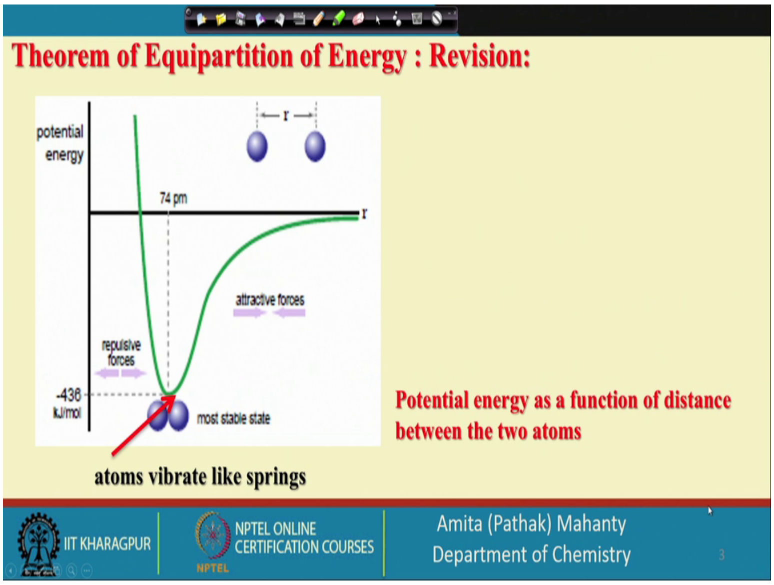

Theorem of Equipartition of Energy : Revision:

Potential energy as a function of distance between the two atoms

atoms vibrate like springs

Amita (Pathak) Mahanty
Department of Chemistry

Usually, what we had talked about in the last class and I could not show you the figure was this is the potential energy diagram which I was talking about, when the attractive forces are dominant and which when the repulsive force is dominant and which is the condition where you have the minimum energy of potential. So, this is the potential energy curve for typical hydrogen atom.

So, if we can go through the previous slides the repulsive forces are supposed to be in contributing to the intermolecular forces when the separation of the molecules are small and the attractive forces are going to be when the intermolecular forces are large. So, this is how we had seen in the potential energy diagram, how we are talking about the internal energy the potential energy being in the repulsive mode or in the attractive mode.

And, then we had seen that at low temperature the average kinetic energy is small and the internal energy is supposed to be corresponding to a bound state where the internal energy is supposed to be negative, when at higher temperature the average kinetic energy of the molecules are high. So, the internal energy is going to be positive and this is going to lead to the molecules having enough energy to escape the intermolecular force to become a gas.

And, what we were talking about in the last one was the minimum potential energy at the point which we are looking here. This is the point where we can a consider the atom as behaving like a spring. So, this is what we had talked about in the last class.

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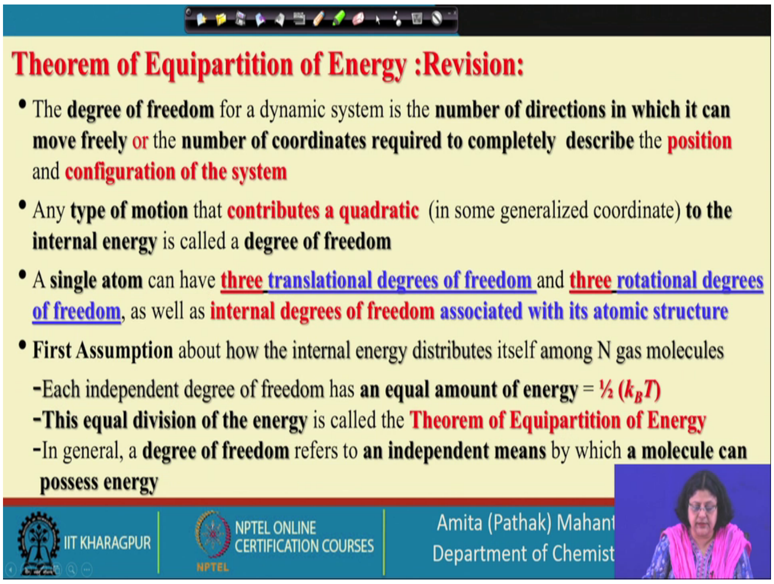

Theorem of Equipartition of Energy :Revision:

- The **degree of freedom** for a dynamic system is the **number of directions in which it can move freely** or the **number of coordinates required to completely describe** the **position** and **configuration of the system**
- Any **type of motion** that **contributes a quadratic** (in some generalized coordinate) **to the internal energy** is called a **degree of freedom**
- A **single atom** can have **three translational degrees of freedom** and **three rotational degrees of freedom**, as well as **internal degrees of freedom associated with its atomic structure**
- **First Assumption** about how the internal energy distributes itself among N gas molecules
  - -Each independent degree of freedom has **an equal amount of energy** = ½ (k_BT)
  - -**This equal division of the energy** is called the **Theorem of Equipartition of Energy**
  - -In general, a **degree of freedom** refers to **an independent means** by which **a molecule can possess energy**

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Department of Chemist

We had talked about what was the degree of freedom. In a dynamic system the number of directions in which a molecule can freely move or you can say the coordinates minimum coordinates required to specify their spatial presence in a particular position or in the space or the configuration of the system is what is known as the degree of freedom.

And, how the energy is distributed in the degrees of freedom will be depending on the type of molecule we are talking about if we have a monatomic molecule or we have a complex if we more than one atom involve a diatomic molecule we had treated that as a complex molecule and in discussed in the last class. What we had done in the assumptions was the first assumption was how the internal energy is distributed among the N number of gas molecules.

And, what we had talked about was each independent degree of freedom has an equal amount of energy of half kT and kT is actually the kB that is a Boltzmann constant we usually say while talking we are and instead of variety is saying kB we all may say kT also.

So, this is the basis of the equipartition energy of energy and based on that we are going to be moving into a new and another concept which is going to be derived from whatever we have discussed in the equipartition theorem.

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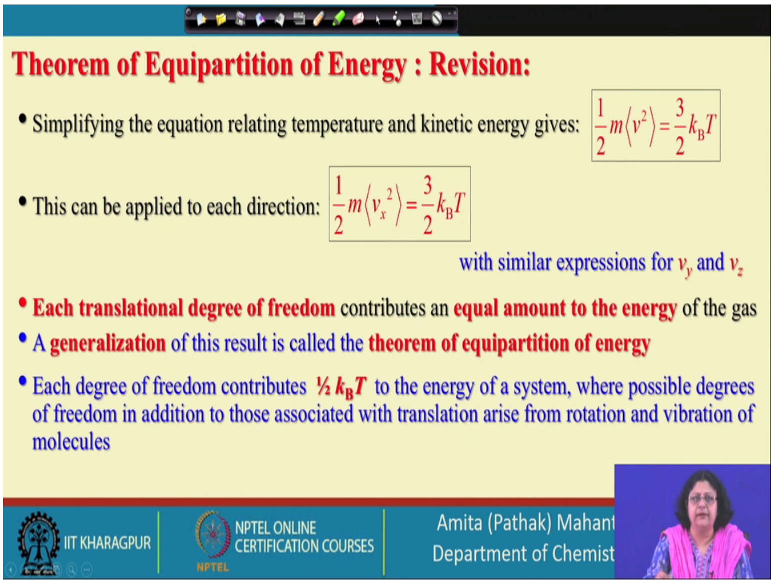

Theorem of Equipartition of Energy : Revision:

- Simplifying the equation relating temperature and kinetic energy gives: $\boxed{\frac{1}{2}m\langle v^2\rangle = \frac{3}{2}k_BT}$
- This can be applied to each direction: $\boxed{\frac{1}{2}m\langle v_x^2\rangle = \frac{3}{2}k_BT}$

with similar expressions for v_y and v_z

- **Each translational degree of freedom** contributes an **equal amount to the energy** of the gas
- A **generalization** of this result is called the **theorem of equipartition of energy**
- Each degree of freedom contributes $\frac{1}{2} k_BT$ to the energy of a system, where possible degrees of freedom in addition to those associated with translation arise from rotation and vibration of molecules

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The equipartition theorem essentially talks about each degree of freedom contributes half kT to the energy of the system wherever, the degree of a freedom is in addition to the associated with the translation or rotational or vibrational motion of the molecule.

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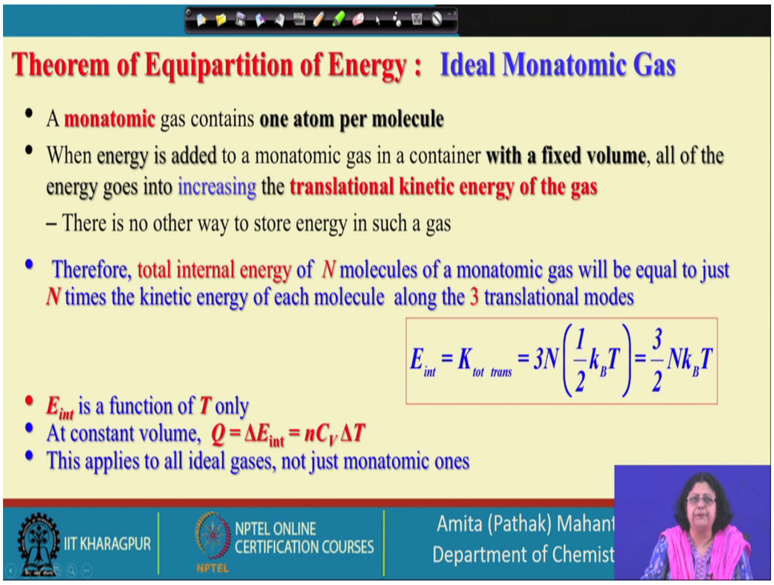

Theorem of Equipartition of Energy : Ideal Monatomic Gas

- A **monatomic** gas contains **one atom per molecule**
- When energy is added to a monatomic gas in a container **with a fixed volume**, all of the energy goes into increasing the **translational kinetic energy of the gas**
  - – There is no other way to store energy in such a gas
- Therefore, total internal energy of N molecules of a monatomic gas will be equal to just N times the kinetic energy of each molecule along the 3 translational modes

$$\boxed{E_{int} = K_{tot\ trans} = 3N\left(\frac{1}{2}k_BT\right) = \frac{3}{2}Nk_BT}$$

- E_{int} is a function of T only
- At constant volume, $Q = \Delta E_{int} = nC_V\Delta T$
- This applies to all ideal gases, not just monatomic ones

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So, this we had talked about in the last class. So, what we are going to do today this is also something which we had discussed in the last class. How the internal energy can be actually for a monatomic gas can be the total translation energy because it is a monatomic gas we have a point atom or a point mass where we do only associate with the translation motion we do not have any other motion associated with it. So, the total internal energy is going to be the measure of the total translational kinetic energy of the system. So, that was as we had if you remember it was taken to be 3 by 2 N k B T, N being the total number of molecules monatomic molecules involve in the system.

So, this was what we had talked about and how internal energy was a function of temperature and how that can be equated to the say specific heat we had talked about in the last class, but will not go into details of it.

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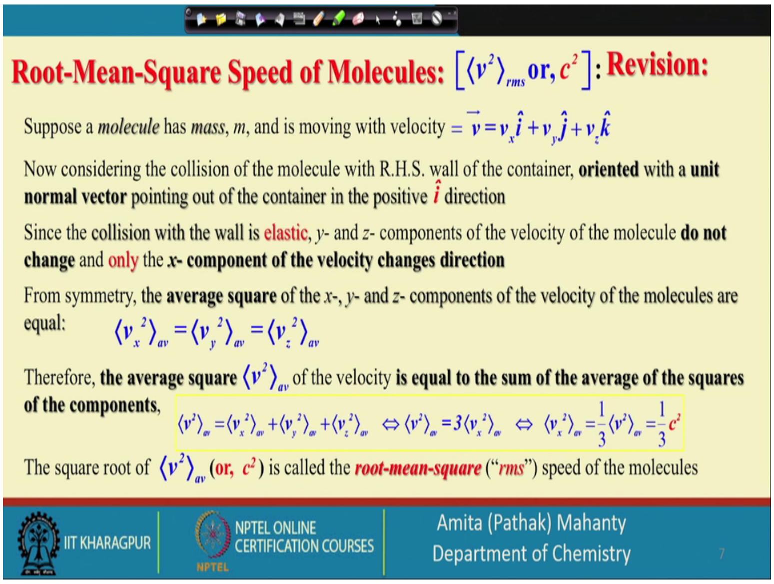

Root-Mean-Square Speed of Molecules: $\left[\langle v^2\rangle_{rms}\text{ or, } c^2\right]$: Revision:

Suppose a *molecule* has *mass*, *m*, and is moving with velocity = $\vec{v} = v_x\hat{i} + v_y\hat{j} + v_z\hat{k}$

Now considering the collision of the molecule with R.H.S. wall of the container, **oriented** with a **unit normal vector** pointing out of the container in the positive $\hat{i}$ direction

Since the collision with the wall is elastic, *y*- and *z*- components of the velocity of the molecule **do not change** and only the ***x*- component of the velocity changes direction**

From symmetry, the **average square** of the *x*-, *y*- and *z*- components of the velocity of the molecules are equal: $\langle v_x^{\,2}\rangle_{av} = \langle v_y^{\,2}\rangle_{av} = \langle v_z^{\,2}\rangle_{av}$

Therefore, **the average square** $\langle v^2\rangle_{av}$ of the velocity **is equal to the sum of the average of the squares of the components,**

$$\boxed{\langle v^2\rangle_{av} = \langle v_x^{\,2}\rangle_{av} + \langle v_y^{\,2}\rangle_{av} + \langle v_z^{\,2}\rangle_{av} \;\Leftrightarrow\; \langle v^2\rangle_{av} = 3\langle v_x^{\,2}\rangle_{av} \;\Leftrightarrow\; \langle v_x^{\,2}\rangle_{av} = \frac{1}{3}\langle v^2\rangle_{av} = \frac{1}{3}c^2}$$

The square root of $\langle v^2\rangle_{av}$ (or, c^2) is called the ***root-mean-square*** ("*rms*") speed of the molecules

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Today we are going to look into another concept which is the root mean square speeds of molecule. I again revise because a root mean square is a term which is very important when we are talking about kinetic energy and I want you people to understand what I am talking about.

See velocity is a vector quantity, suppose I have a mass m and that is moving with a velocity v, so, it will have since it is a spatial x, y, 3 coordinates are there in space x, y, z. So, if I say that the velocity is v x along the I axis and along x axis v v y along the y axis and v z along the z axis.

And, here what I am considering the molecules and the collisions of the molecules if it is happening on the right hand side of the container then the orientation which is going to what we were looking into is a unit normal vector unit normal vector if I am talking about right hand side I am talking about the x axis, then the unit normal vector pointing on towards the container is going to be the positive i direction. So, this is the I factor which is associated with the vector.

So, since it is a elastic collision which we are talking about. So, if I am talking about the direction only x then the I will consider that the z and y axis components are remaining unchanged, but that concept is that the molecules are going to be probably oriented everywhere not only in x axis, y axis, as well as z axis when I am talking about the collision along the right hand side when I have the unit normal vector as i positive i then I am talking about a positive x direction.

Now, similarly, if I talk about the symmetry because it is equally probable that you have the molecules in the x and y and z axis so, what I can say this in the average square of the x and the velocities in the x component, y component and the z component can be given as something like this; v x square v equal to v z v y square equal to v z square this average energy along the xyz direction that is going to be equally probable and there should be equal because we are talking about in terms of the symmetry because they are all equally probable in the x, y, z direction.

So, the average square can be given as the average of this is equal to the sum of the average of the squares of the components. So, I can write down the sum of the squares of the component and this is since these are all equally probable so, I can write down this in terms of x, y or z I can write down 3 x as the average velocity square is 3 x in the x direction or I can say the average along the x axis 1 by 3 whatever the term which I have taken this was actually derived from the Pythagoras theorem, if you remember in the previous classes we talked about 1 by 3.

This velocity square is what we are talking about is the root mean square and this sometimes is referred to as c in some of the textbook. So, please do not confuse the this is actually the average velocities which we are looking into, but we are saying this is as root mean square velocity because we are going to be taking the exact meaning of the term. We are going to take the under root of the mean of the squares of the velocities of the molecules.

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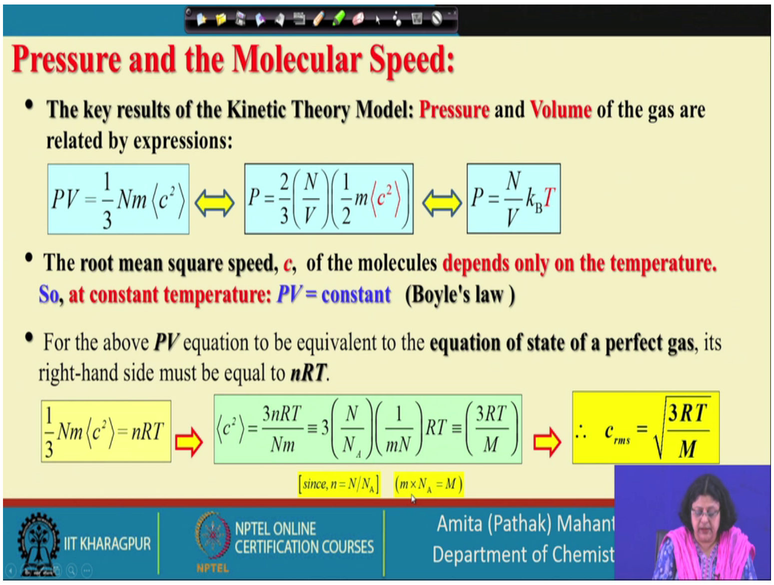

Pressure and the Molecular Speed:

- **The key results of the Kinetic Theory Model: Pressure and Volume of the gas are related by expressions:**

$$\boxed{PV = \frac{1}{3}Nm\langle c^2\rangle} \Leftrightarrow \boxed{P = \frac{2}{3}\left(\frac{N}{V}\right)\left(\frac{1}{2}m\langle c^2\rangle\right)} \Leftrightarrow \boxed{P = \frac{N}{V}k_BT}$$

- **The root mean square speed, c, of the molecules depends only on the temperature. So, at constant temperature: PV = constant (Boyle's law)**
- For the above ***PV*** equation to be equivalent to the **equation of state of a perfect gas**, its right-hand side must be equal to ***nRT***.

$$\boxed{\frac{1}{3}Nm\langle c^2\rangle = nRT} \Rightarrow \boxed{\langle c^2\rangle = \frac{3nRT}{Nm} \equiv 3\left(\frac{N}{N_A}\right)\left(\frac{1}{mN}\right)RT \equiv \left(\frac{3RT}{M}\right)} \Rightarrow \boxed{\therefore\ c_{rms} = \sqrt{\frac{3RT}{M}}}$$

$[\text{since}, n = N/N_A]$ $(m \times N_A = M)$

So, this is how we are going to define the rms velocity and according to the previous derivations which we had carried out the key results of the kinetic model was the pressure volume of the gas could be related by this expression. I am not going to read out the expression it is obviously, visible to you PV equal to 1 by 3 nm into this, this c is the root mean square velocity I am talking about this is also the root mean square velocity again this is related to the k B of the Boltzmann constant to temperature.

So, we can see that pressure is directly proportional to the root mean square velocity. The root mean square velocity c of the molecule depends on the temperature from this expression if we can equate then we can see that the root mean square velocity is directly dependent on the temperature and if it the temperature is constant then from this expression we can say PV if you multiply and temperature is a constant then for a N number of molecules in a system the PV becomes a constant. So, that is something which is the Boyle’s law we are going to discuss that the verifications of the gas laws a while from now.

So, again if I have the PV equation and I say that equation it can be equivalent or can be equated to a equation of state of a perfect gas that is PV equal to nRT then this part PV should be equal to nRT. So, that is how I equated this P this is also PV. So, I equated to nRT from there what I am a trying to find out is the root means expression for the root mean square velocity. Root mean square velocity is equal to is going to be given by from here you can understand it is going to be nRT divided by m into capital N into in the numerator you have the 3.

So, what you are doing here you just see what you are putting your substituting for n the number of moles is going to be total number of molecules by the Avogadro number and you know that small mass is the mass of the gas into the Avogadro number gives you the molar mass of the system. So, if you are substitute this in place of the M then you get the expression this is the expression of root mean square velocity.

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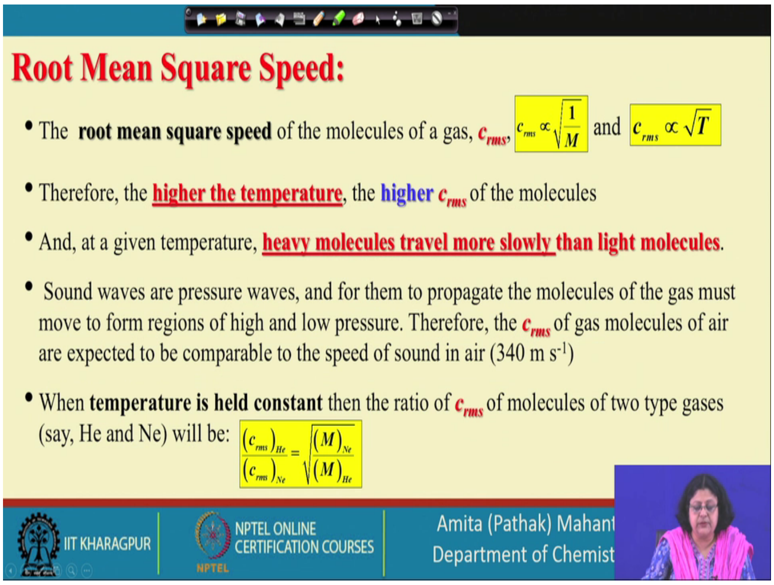

Root Mean Square Speed:

- The **root mean square speed** of the molecules of a gas, c_{rms}, $c_{rms} \propto \sqrt{\frac{1}{M}}$ and $c_{rms} \propto \sqrt{T}$
- Therefore, the **higher the temperature**, the **higher c_{rms}** of the molecules
- And, at a given temperature, **heavy molecules travel more slowly than light molecules**.
- Sound waves are pressure waves, and for them to propagate the molecules of the gas must move to form regions of high and low pressure. Therefore, the c_{rms} of gas molecules of air are expected to be comparable to the speed of sound in air (340 m s^{-1})
- When **temperature is held constant** then the ratio of c_{rms} of molecules of two type gases (say, He and Ne) will be: $\frac{(c_{rms})_{He}}{(c_{rms})_{Ne}} = \sqrt{\frac{(M)_{Ne}}{(M)_{He}}}$

So, what is this root mean square velocity from the expression it is clear that it is going to be directly proportional to the temperature and it is inversely proportional to directly proportional to under root of temperature and it is inversely proportional to under root of the molar mass. So, the higher the temperature will be you will get the higher value of the rms the root mean squared speed of the molecule. For a given temperature; that means, if you since it is the rms velocity is inversely proportional to the mass.

So, the heavier the molecule will be slower that with movement of them particle will be because the molar mass will be higher. So, it is since it is inversely proportional. So, the speed will be lower. So, heavier molecules will move slowly compared to the lighter ones. Again, I one important factor is sound waves are pressure waves, ok. So, these propagate as they propagate the molecules of the gas must move from a higher pressure region to a lower pressure region.

So, since it is propagating through the air so, we can say that the rms speed of the molecules of air is expected to be comparable up to the speed of sound in air that is around 340 meters per second. Now, again, if you see the expressions expression of the root mean square velocity, you will see that if you hold the temperature constant for a system if you have two system, two type of gases if you hold the system at the same temperature then you will see the ratios of the rms velocities will be inversely proportional to the under root of the masses.

So, if you see here if you have helium, the rms velocity of helium and rms velocity of neon if you can take the ratios it will be under root of the molar mass of neon divided by the molar mass of the helium. So, these are things which is important when you are talking doing calculations or numericals.

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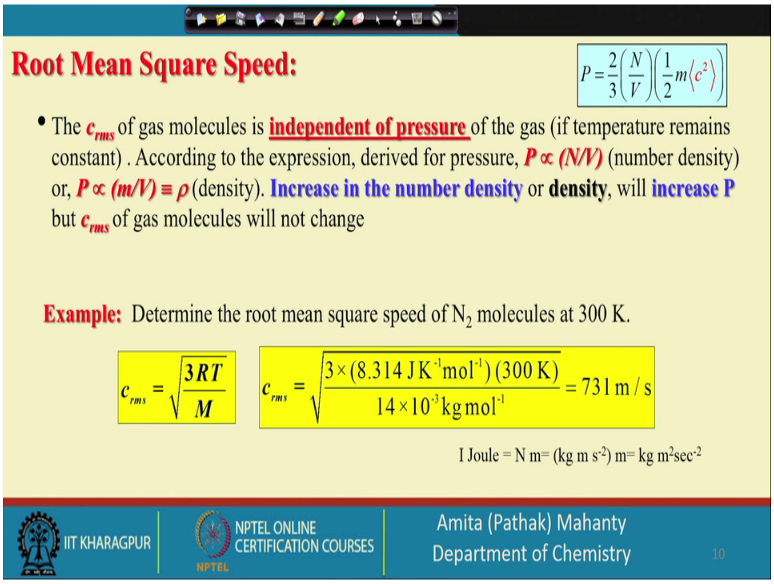

Root Mean Square Speed:

$$P = \frac{2}{3}\left(\frac{N}{V}\right)\left(\frac{1}{2}m\langle c^2\rangle\right)$$

- The c_{rms} of gas molecules is **independent of pressure** of the gas (if temperature remains constant) . According to the expression, derived for pressure, $P \propto (N/V)$ (number density) or, $P \propto (m/V) \equiv \rho$ (density). **Increase in the number density** or **density**, will **increase P** but c_{rms} of gas molecules will not change

Example: Determine the root mean square speed of N_2 molecules at 300 K.

$$c_{rms} = \sqrt{\frac{3RT}{M}}$$

$$c_{rms} = \sqrt{\frac{3 \times (8.314\ \mathrm{J K^{-1} mol^{-1}})\ (300\ \mathrm{K})}{14 \times 10^{-3}\ \mathrm{kg\ mol^{-1}}}} = 731\ \mathrm{m/s}$$

I Joule = N m= (kg m s^{-2}) m= kg m^{2}sec^{-2}

So, another factor which needs to be remembered we have this expression of pressure. So, you will see from in the expression of the rms speed we do not have a term of pressure. So, rms speed of a gas molecule is independent of the pressure. If you if the temperature is held constant then the molecules the gas molecules are independent of the pressure of the gas. If you see the expression of the pressure what you see is pressure is directly proportional to the N by V. What is N by V? It is the number of molecules per unit volume of the gas that is the container which we are talking about this V. So, this is the number density.

Similarly, if you can see the in the expression you have m, small m it is a number of mass of the gas. So, mass of the gas divided by the total volume will give you the density. So, as the number density or the density of the gas increases, what happens, the pressure increases, but it does not have any effect on the rms speed of the molecules, they remain unchanged.

Now, let us see one example this is a solved example for finding out the rms speed. rms speed is the expression which we have here is 3 RT by M. So, if you want to find out the rms speed for N nitrogen molecules at 300 degree Kelvin then you can put the values here as you see 3 into you put the values the gas constant 8.314 joules to the Kelvin inverse mol inverse into the temperature which is in Kelvin. And here if you remember you have to convert the molar mass into kg's per mole because you have in joules. Joules is 1 joule is what is given by Newton meter and that is if you convert that in into kilo meters per second square that will be again into multiplied by m.

So, if you see the kg's will get cancelled. So, what you are left with meter square per second square. So, if you take the under root of that you get by in the units of meters per second. So, this is one of the a solved examples you can see for finding out the root mean square speed of nitrogen molecule.

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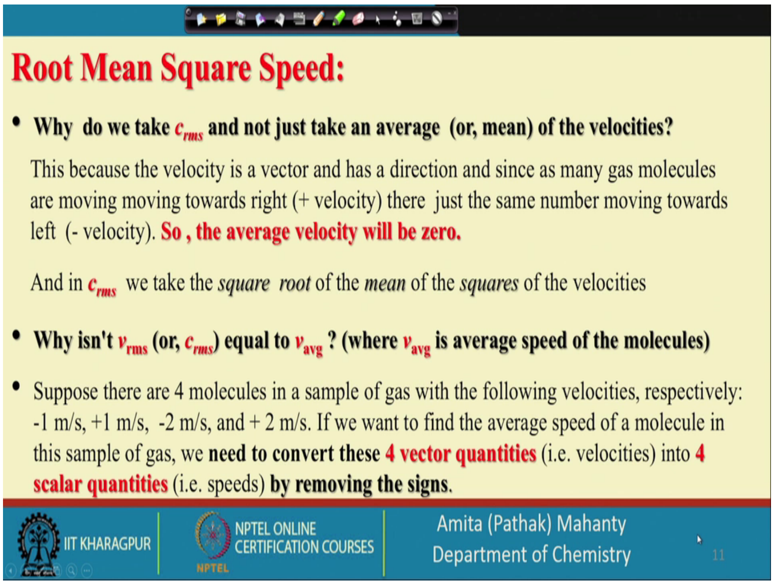

Root Mean Square Speed:

- **Why do we take c_{rms} and not just take an average (or, mean) of the velocities?**

  This because the velocity is a vector and has a direction and since as many gas molecules are moving moving towards right (+ velocity) there just the same number moving towards left (- velocity). **So , the average velocity will be zero.**

  And in c_{rms} we take the *square root* of the *mean* of the *squares* of the velocities

- **Why isn't v_{rms} (or, c_{rms}) equal to v_{avg} ? (where v_{avg} is average speed of the molecules)**
- Suppose there are 4 molecules in a sample of gas with the following velocities, respectively: -1 m/s, +1 m/s, -2 m/s, and + 2 m/s. If we want to find the average speed of a molecule in this sample of gas, we **need to convert these 4 vector quantities** (i.e. velocities) into **4 scalar quantities** (i.e. speeds) **by removing the signs.**

So, why do we take the root mean square velocities, why cannot we just take the average velocities? See, why cannot we take average velocities, see the molecules are moving along the x axis, y axis and z axis we are looking into the movement of the molecules along the three axises.

So, it is not an adjust the x axis, but it is in the positive x axis and the negative x axis. Since the according to the symmetry rule it is possible that all the molecules are going to be equally likely to move in the x plus direction and in the x minus direction. So, what if even if you take the mean what happens you add upcoming to 0, because minus velocity and plus velocity is well eventually come to giving you a total average of 0.

So, that is the reason we do not take the mean velocities, but we go for taking the root mean square velocity and the name itself suggests what the root mean square velocity is. It is the square root of the mean of the squares of the velocities. So, what is the difference between having a rms speed and the average speed, why do we take it? Let us see examples out here. Suppose, you have 4 molecules of gas moving with the speed minus 1 meters per second, plus 1 meters per second, minus 2 meters per second, plus 2 meters per second.

So, if you want to find out the average speed of this molecule we since these are vector quantities and they have positive and negative terms, we have to first convert these 4 vectors into a scalar component to spine the speed. So, the we can you do that by removing the signs.

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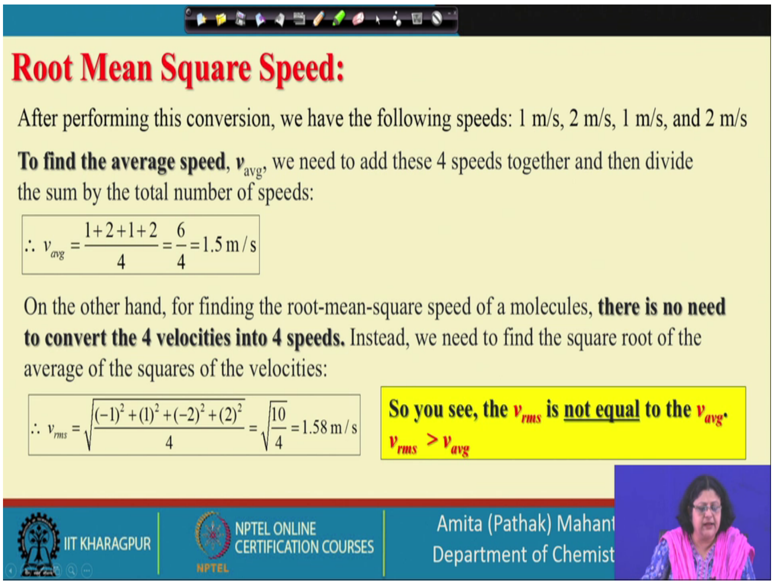

Root Mean Square Speed:

After performing this conversion, we have the following speeds: 1 m/s, 2 m/s, 1 m/s, and 2 m/s

To find the average speed, v_{avg}, we need to add these 4 speeds together and then divide the sum by the total number of speeds:

$$\therefore v_{avg} = \frac{1+2+1+2}{4} = \frac{6}{4} = 1.5\,\mathrm{m/s}$$

On the other hand, for finding the root-mean-square speed of a molecules, **there is no need to convert the 4 velocities into 4 speeds.** Instead, we need to find the square root of the average of the squares of the velocities:

$$\therefore v_{rms} = \sqrt{\frac{(-1)^2+(1)^2+(-2)^2+(2)^2}{4}} = \sqrt{\frac{10}{4}} = 1.58\,\mathrm{m/s}$$

So you see, the v_{rms} is <u>not equal</u> to the v_{avg}.

$v_{rms} > v_{avg}$

So, if you remove the signs what you are doing of craft upon doing the conversion you get all the positive terms. You remove the signs to get the scalar quantity of speed, instead of talking about velocity we are talking about speed now. So, the average speed will be what we add the all the speed values 1 plus 2 plus 1 plus 2 divided by 4, the total number of molecules you get 1.5 meters per second.

Now, on the other hand how do you find out the root mean square velocities? Here we do not need to convert the vectors to scalars because you have the square terms as soon because we are taking this mean of the squares. So, we are squaring the each term. So, the negative terms will be automatically becoming positive we do not need to look into the conversion of a vector to scalar.

So, that is the reason why we use rms speed. So, minus 1 square plus 1 square minus 2 square; So, it becomes 10 divided by 4, you take under root, you get 1.5. Now, you see if you take the rms speed and the average speed if you while calculating by converting these vectors into the scalar quantity in spite of that you will see that the rms speed is 1.58 and here it is 1.5. So, the rms speed is going to be higher and then the average speed and they are not equal. So, the root mean square of speed is not exactly equal to the average speed and that is that is why we always concentrate on the rms speed because we are dealing actually with the vectors species of velocity.

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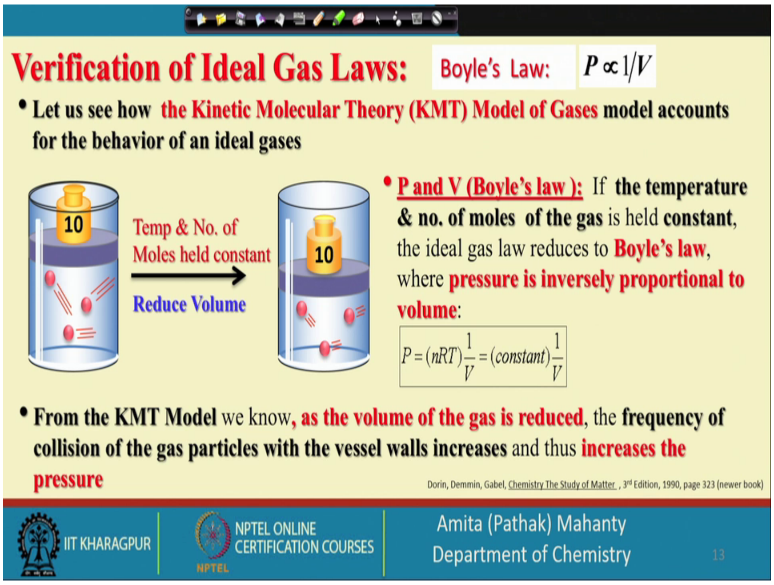

Verification of Ideal Gas Laws:

Boyle's Law: $P \propto 1/V$

- **Let us see how the Kinetic Molecular Theory (KMT) Model of Gases model accounts for the behavior of an ideal gases**

- **P and V (Boyle's law):** If **the temperature & no. of moles of the gas** is held **constant**, the ideal gas law reduces to **Boyle's law**, where **pressure is inversely proportional to volume**:

$$P = (nRT)\frac{1}{V} = (constant)\frac{1}{V}$$

- **From the KMT Model** we know, **as the volume of the gas is reduced**, the **frequency of collision of the gas particles with the vessel walls increases** and thus **increases the pressure**

Dorin, Demmin, Gabel, Chemistry The Study of Matter , 3rd Edition, 1990, page 323 (newer book)

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Now, today we are going to be looking into some of the terms which we have dealt in the beginning we have talked about the gas laws. Now, we will do the verification of gas law based on whatever we have found from the kinetic molecular theory of gases. So, usually the kinetic molecular theory is in various books you will find that it is abbreviated as KMT, KMT model of gas.

So, now we are going to account of the take care of the whatever we have observed in our kinetic model is it going to be valid for the gas equations which we said is going to be it should be valid according to our findings. So, what we talk about we talk about pressure and volume conditions first. If the temperature and number of moles of the gas is held constant the temperature and the number of moles is held constant then according to Boyle's law we know the pressure is inversely proportional to the volume, right. This is going to be the expression of inert gas law if the temperature is held constant and the number of moles are held constant, this part is constant. So, pressure is going to be inversely proportional to the volume.

Now, how does the kinetic theory justify this? When you have a volume of the gas reduced, what happens when you reduce the volume of the gas from here to here, what happens is the space you are providing for the molecules to move is confined further. So, there is the number of collisions which is going to be taking place the frequency of the collisions which is going to take place with the walls of the container is going to increase and what is pressure? Pressure is that the is proportional to the number of free the collisions the gas molecules are making on the surface. So, when you are reducing the volume then what is likely the number of the frequency of the collisions are with the walls are likely to increase what will it give rise to? It will give rise to the increase in pressure.

So, according to the kinetic theory of gases if you are in decreasing the volume of the gas then the pressure of the gas is supposed to rise because of the increasing number of collisions with the vessel surface, continuous surface. So, that is nothing, but the Boyle's law. So, we have proved one of the gas laws.

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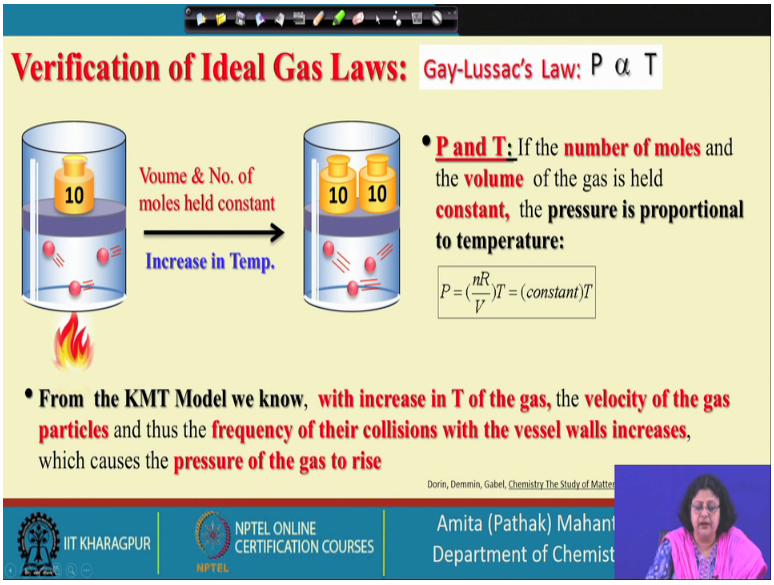

Verification of Ideal Gas Laws: Gay-Lussac's Law: P α T

- **P and T:** If the **number of moles** and the **volume** of the gas is held **constant,** the **pressure is proportional to temperature:**

$$P = \left(\frac{nR}{V}\right)T = (constant)T$$

- **From the KMT Model we know, with increase in T of the gas,** the **velocity of the gas particles** and thus the **frequency of their collisions with the vessel walls increases,** which causes the **pressure of the gas to rise**

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Department of Chemist

And, the next one is suppose the number of moles, number of moles and the volume is held constant then what happens the pressure is directly proportional to the temperature. This is what? This is the Gay Lussac's law. Gay Lussac's law says pressure is directly proportional to temperature.

Now, what happens is through the kinetic model do we justify what we have gone through the gas law model? Now, what we have we have number of moles fixed we have volume fixed we are not changing the volume. Now, I increase the temperature by heating supplying some amount of heat to the container. What happens as you rise the temperature, what happens the velocity of the gases will increase because the temperature is direct measure of the kinetic energy. So, the kinetic energy is increasing to a number of the speed of the molecules is going to increase the velocity of the molecules are increment to increase and so will be the frequency of the collision with the walls.

So, as you increase the temperature the frequency of the collisions with the wall also increases because now the speed of the molecule has increase, they are moving much faster. So, there in the rate of collision with the walls is going to be increasing the frequency of collision is increasing and if the impact which is we get from the collision is actually the measure of the pressure. So, the pressure is going to rise. So, what we had gone from the Gay-Lussac's law as pressure being directly proportional to temperature is what we are going to see in this kinetic model as well.

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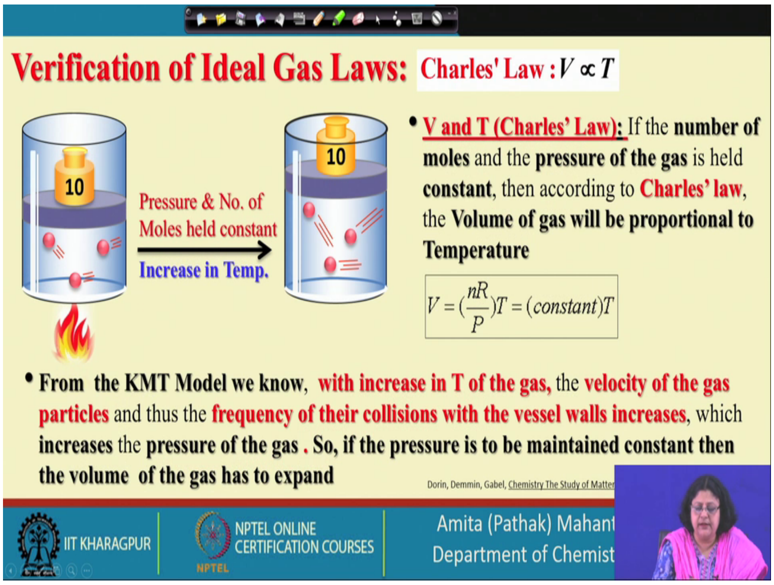

Verification of Ideal Gas Laws: Charles' Law : $V \propto T$

- **V and T (Charles' Law):** If the **number of moles** and the **pressure of the gas** is held **constant**, then according to **Charles' law**, the **Volume of gas will be proportional to Temperature**

$$V = \left(\frac{nR}{P}\right)T = (constant)T$$

- **From the KMT Model we know, with increase in T of the gas,** the **velocity of the gas particles** and thus the **frequency of their collisions with the vessel walls increases**, which **increases** the **pressure of the gas . So, if the pressure is to be maintained constant then the volume of the gas has to expand**

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Now, the next one is the Charles' law. What are the Charles' law say? Number of moles are held constant, the substance is constant and the pressure of the gas is held constant. So, pressure is held constant number of moles are held constant, what we see is Charles' law. According to Charles' the volume is going to be proportional to the temperature. Now, what we are doing here? We are taking the number of moles and the pressure as constant. So, if I have a weight I have equal weights on the top; that means, the pressure I am trying to maintain as constant which is pressing the gas.

So, the pressure maintained constant if I start heating the gas container. So, what happens the gas can the gas will increase in kinetic energy, they will be moving much faster. So, they will start hitting the walls of the container, much faster the frequency of collisions will be much faster. So, in the under that condition what you have you have the increase in pressure, but we are having a condition where we want to maintain the pressure constant.

So, to maintain the pressure constant what we have to do, we have to expand the gas, right. So, the gas expands and you have a increase in volume. So, as we increase the temperature we actually need to increase the volume of the gas to keep the temperature constant for a fixed amount of gas. This is the Charles' law.

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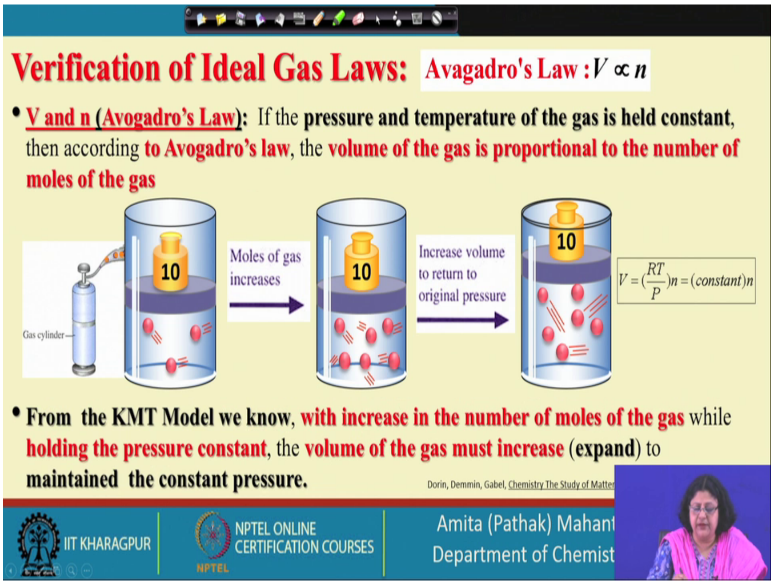

Verification of Ideal Gas Laws: Avagadro's Law : $V \propto n$

- **V and n (Avogadro's Law):** If the **pressure and temperature of the gas is held constant**, then according **to Avogadro's law**, the **volume of the gas is proportional to the number of moles of the gas**

$V = (\frac{RT}{P})n = (constant)n$

- **From the KMT Model we know, with increase in the number of moles of the gas** while **holding the pressure constant**, the **volume of the gas must increase** (**expand**) to **maintained the constant pressure.**

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And, then we have the Avogadro's law. Avogadro's law what does it say if the pressure and temperature gas is held constant then the according to Avogadro's law so, the volume of the gas is proportional to the number of moles of the gas, ok.

So, suppose I have a gas container, where I have the pressure of each of the gas being constant. Now, I what I do I have a separate a external gas cylinder supply I attach the gas molecules in the cylinder to the container and pass the gas molecule. Now, I am supplying the increased number of moles. If I am increasing number of moles what happens it is going into this with increasing number of moles of the gas and we are holding the pressure trying to hold the pressure constant what is going to be happening in the number of collisions are going to increase. So, the pressure is supposed to be increasing, but we are keeping the pressure constant. To maintain the pressure constant, what we have to do? We have to expand the gas.

So, if we are increasing the number of molecules or number of moles of the gas, but you have the conditions of constant pressure then the volume of the gas should be proportional to the number of moles. This is also verified by the kinetic model.

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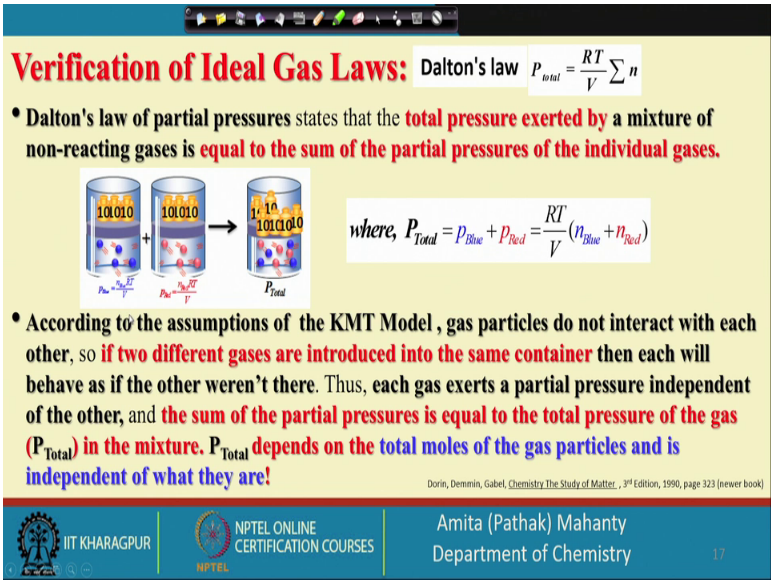

Verification of Ideal Gas Laws: Dalton's law $P_{total} = \frac{RT}{V}\sum n$

- **Dalton's law of partial pressures** states that the **total pressure exerted by a mixture of non-reacting gases is equal to the sum of the partial pressures of the individual gases.**

where, $P_{Total} = p_{Blue} + p_{Red} = \frac{RT}{V}(n_{Blue} + n_{Red})$

- **According to the assumptions of the KMT Model, gas particles do not interact with each other**, so **if two different gases are introduced into the same container then each will behave as if the other weren't there**. Thus, **each gas exerts a partial pressure independent of the other,** and **the sum of the partial pressures is equal to the total pressure of the gas (P_{Total}) in the mixture. P_{Total} depends on the total moles of the gas particles and is independent of what they are!**

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Then last one is the verification of ideal gas law, which is the Dalton's partial pressure law. Dalton's partial pressure law states the total pressure exerted by a mixture of non reacting gases is equal to the sum of the partial pressure of the individual gases.

So, what I have, I suppose have two types of gases; one corresponding of blue colored balls and the other color red colored balls. What happens, the pressure exerted by each; what is going to be the pressure exerted by this is going to be nRT by V, ok. Here nRT is the number of moles for the blue balls, here it is the nRT for the number of red balls. So, what I have, the total pressure will be since they are not reacting according to the kinetic model we do not have interaction between the molecules.

So, what happens the total pressure is now going to be the partial pressure of blue in the blue system as well as the red system. When we mix these two they are not reacting gases. So, they will behave as if the two different gases are introduced in the same container will each have behaved like the other is not there.

The partial pressure of the gas is as if only one gas is the existing the other is not existing at all. So, they are not dependent on the other gas. So, each gas exerts a partial pressure independent of each other. The sum of the partial pressures if I take the sum of the partial pressures this is I can take the number of RT by V as common and then what I get is the number of moles of the blue and the number of moles of the red.

So, what I get is the expression of Dalton's pressure a partial pressure and law of partial pressure that is P equal to total pressure equal to RT by V summation of the number of moles. So, what we have, the total pressure depends on the total number of moles of the gas particles and it is independent of the type of the gas which we are talking about. So, the total pressure is going to be the sum of the partial pressures, because they are not reacting gases. And, according to kinetic theory of gases the molecules do not have any interaction with each other. So, they are going to be behaving independently.

So, if you are talking about a gas pressure here we are talking about one type of gas pressure contributed by this system and now, that is contributed by this system that the total pressure of this system and this system together should be the pressure. So, this is the pressure partial pressure which we call and this is the partial pressure we call for this system, the total will be the summation of the partial pressures.

So, this is what I have discussed. Next class we are going to be talking about the molecular distribution of speed.

Thank you so much.