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Liquids are incompressible; gases are not incompressible.

When you decrease the volume of a gas by 50%, that's no problem.

It's impossible to do that for a liquid.

In liquids, the atoms and the molecules effectively touch each other, whereas in gases, they are very far apart, so that's why you can compress the gases.

If you take air at one atmospheres, the density is a thousand times less than the density of water.

What it tells you is that the molecules are much further apart.

It is an experimental fact that there is a simple relation between the pressure that you see there, the volume of a gas, the temperature of a gas in degrees Kelvin, and the number of molecules that you have.

Now, when you see the word "molecules," I may often mean "atoms." I realize that helium and neon and krypton and argon are atomic gases, and that O_2 and H_2 and CO_2 are molecular gases.

So I will use that word "molecules" even when I mean "atoms," and maybe vice versa, just for simplicity.

The relation that exists between these quantities, PV equals nRT: pressure, volume, n is the number of moles--

I'll get back to that--

R is the universal gas constant, which is 8.3 joules per degree Kelvin, and T must be in degrees Kelvin.

So, what is a mole? A mole has always about 6.02 times ten to the 23 molecules, or atoms, in the case that you have helium, but I will call that molecules.

And this number is called Avogadro's number.

So that's the definition of a mole.

If you take a mole of helium, or a mole of oxygen, or CO_2, or N_2, it doesn't matter, it always has this number of molecules, approximately.

Now, each of these substances have very different masses.

If I take, for instance, carbon, then one mole of carbon would weigh very close to 12 grams.

If I take helium, one mole of helium would weigh very close to four grams.

And if I took oxygen two, O2, then one mole would be very close to 32 grams.

So the masses are very different in a mole but not the number of molecules or the number of atoms.

When I take a neutral atom, then we have a nucleus, and the nucleus contains protons and neutrons.

It has Z protons and it has N neutrons.

The protons are positively charged, and it has Z electrons if it is a neutral atom.

There is almost no weight in the electrons; you can almost ignore that.

Everything is in the protons and in the neutrons.

N plus Z is called A, and that's called the atomic mass number.

Let's look at carbon in a little bit more detail.

If we have carbon--

and I call it carbon 12 for now, you'll see shortly why--

then carbon has always six protons in the nucleus; otherwise it isn't carbon.

And when it has six neutrons, then A is 12.

That's why we call it carbon 12.

So the atomic mass number of carbon is 12, but if you had, for instance, carbon 14--

which happens to be radioactive--

again, six protons, otherwise it wouldn't be carbon, you would have eight neutrons now, and now you would have...

atomic mass number would be 14.

A mole is this number in grams, and so you see carbon...

is the atomic mass number in grams--

you see 12 there.

If you go to helium, it has two protons and two neutrons, so A is four--

that's why you see your four grams.

If you take oxygen, it has eight protons and eight neutrons, so A is 16, but you have O2 in gas form, so now your atomic mass number has to be doubled to 32.

And so a mole of O2 is therefore 32 grams.

In fact, Avogadro's number is defined through carbon 12.

If you take 12 grams of carbon 12, and you count the number of atoms that you have, then you find exactly Avogadro's number.

That's the definition of that number, and that's very close to what we have there, 6.02 times ten to the 23rd.

The mass of the proton and the mass of the neutron are nearly equal.

I wrote down m2 for the mass of the neutron; of course, that should have been m of n.

So the mass of a molecule, or an atom, whatever the case may be, would be this number A--

because that's the sum of the protons and neutrons--

times the mass of the neutrons and the protons.

And so this is A times--

approximately, I should put a wiggle here--

1.66 times ten to the minus 27 kilograms.

So that's now an individual mass of either an atom or a molecule, and all that information, you have there and that's, of course, on the Web.

So, let's do a trivial example.

I take gases, any kind of gas--

you choose whatever you want--

and I take one atmosphere.

So that means that the pressure is 1.03 times ten to the fifth pascal.

I do it at room temperature, so T is 293 degrees Kelvin.

And I take in all cases only one mole, so n is one.

And I'm asking you now, what will be the volume of that gas? Well, you take the gas law, and it tells you that V, the volume, equals nRT divided by P.

You know n is one.

You know R, 1.03...

excuse me, you know...

[laughs]: I'm a little bit ahead of myself.

You know R, which is 8.3, you know the temperature, which is 293, and you know the pressure, which is 1.03 times ten to the fifth.

And when you calculate that, you find something very close to 24 liters, and a liter is about a thousand cubic centimeters.

And it's independent of whether it's helium or oxygen or nitrogen or CO2.

As long as you have a gas, one mole at one atmosphere pressure and room temperature always has the same volume of about 24 liters.

If a gas obeys that law exactly, we call it an ideal gas.

That's why we call that the ideal-gas law.

And many gases come very close to that.

In fact, if you took oxygen, O2, and you take one mole of oxygen at room temperature and at one atmosphere pressure and you were to calculate its volume, the actual volume that you measure is only one-tenth of a percent smaller than what you would have found with the ideal-gas law.

If you do it at 20 atmospheres, it would still be only two percent smaller, so it's a very good approximation in many cases.

What is very surprising, that in this ideal-gas law, the mass of the atoms and the molecules do not show up at all.

And that is very puzzling--

you wouldn't expect that at all.

And I'll show you why you wouldn't expect that.

Let's take two different kinds of gases with very different masses of the molecules, but we have the same number of moles, we have the same volume, we have the same temperature and therefore, we must have the same pressure, according to the ideal-gas law.

But the masses of the molecules--

very different.

So, here we have some of these molecules
and the number of... density is the same, because the number of atoms is the same and the volume is the same.

Now, these molecules are flying in all directions with different speeds.

I will just now, for simplicity, take some average speed, and I assume this is going in this direction.

It's heading for the wall of the container, this area.

It hits the wall, there's an elastic collision, and it comes back in exactly the same direction.

So there is momentum transfer, and the momentum transfer for one collision is 2mv, because it comes in with mv in this direction, it comes back with mv in that direction, so the momentum transfer is 2mv.

But I'm interested in the momentum transfer per second, not just for one molecule.

And now, of course, I have to multiply by the velocity, because if the velocity is high, you have a lot of bombardments per second on here.

For each bombardment, this is the momentum transfer, but if there are many, well, you have to multiply that, of course, then, by the speed.

So the momentum transfer per second is proportional, let's say, to mv squared.

mv comes from the momentum, from one particle, and v comes from the fact that...

the number that hit it per second.

Now, momentum transfer per second is clearly...

It's a force, proportional to the force, and that is proportional to the pressure.

And yet the pressure is not affected by the mass, notice? If these are the same, the pressure must also be the same.

And so there's only one conclusion that you can draw, which is very nonintuitive--

that the pressure can only be the same if, for a given temperature, this product, mv squared, is independent of the mass of the molecule.

How can mv squared possibly be independent of the mass of the molecule? There's only one way that that's possible--

that if you take two different masses, two gases with totally different mass of the molecules, that this product is always the same for a given temperature, and that, indeed, is the case.

So if you take, for instance, helium, and you compare that with oxygen and the ratio of these two masses is four to 32--

this is eight times heavier than this one--

and you have them at a certain temperature, then the mass of the helium times the mean speed of the helium squared is the mass of the oxygen molecules times the mean speed squared of the oxygen.

And so what that means is that if the oxygen molecule is eight times more massive--

this one--

then the velocity is the square root of eight times smaller, because the ratio, 32 over four, is eight.

So the oxygen molecules have a lower speed, so that the product, mv squared, is always the same.

Oxygen molecules at room temperatures have a speed of about 480 meters per second, and so the helium is the square root of eight times higher.

If you have a mixture of oxygen and helium, then one gas would have an average speed of 480 meters per second for the molecules, and the other would be 1,350 meters per second.

That is the only way that the gas law can hold.

It is a consequence of the gas law.

You very often see the gas law written in a different way, and you see it written as PV equals capital N times k times T.

You see that also there.

And this N, now, is the total number of molecules that you have.

Don't confuse that with little n, which is the number of moles that you have.

So this is the total number of molecules,
and k is called Boltzmann's constant.

And since Nk must be the same as nR, if you compare the two gas laws, the one up there and the one here, you can see that the way that k is defined, it is nothing but R divided by N of A.

k is R divided by Avogadro's number, because little n, remember, is the total number of molecules that you have divided by N of A.

So if you substitute that, you get this as a result, and so this is 8.3 divided by 6.02 times ten to the 23rd, and that is about 1.38 times ten to the minus 23rd joules per Kelvin.

And you can use that number.

If you want to, you can use either this relationship or you can use that one, whichever is convenient.

They are identical.

So, let's now bring the ideal-gas law to a test.

And the way I'm going to bring it to a test is as follows: I have here a volume, copper.

There's air inside.

And I'm going to heat that up.

I'll make you a drawing on the blackboard.

Here is that volume, and there's an extremely thin tube, which has almost no volume, and at the end is a pressure gauge.

So there is a pressure gauge here which gives us the pressure, and the pressure is given in pounds per square inch, much to my regret, but that's the way it is.

For your recollection, one atmosphere is approximately 15 pounds per square inch.

It is also a gauge that measures overpressure.

In other words, if you expose it to one atmosphere, it will read zero.

Just when you go to the gas station and you measure the pressure of your tires, that is also a gauge that measures overpressure, the difference between inside and outside.

This has a valve here, and I can connect it with the outside world.

And when I do that, then the pressure is very close to one atmosphere, regardless of what the temperature is here, because it's connected to the universe, so the pressure in here will then be one atmosphere.

And that's the situation that we have now when this valve is open, and it is in melting ice, And so melting ice--

T1 is 273 degrees Kelvin.

P1 is one atmosphere, and V1 is some value that I don't know.

I close this valve.

Thereby the number of moles of air in there is fixed; that's not going to change.

So we have one atmosphere inside now,
and whatever n is, it's of no importance, but it's not going to change.

Now I'm going to put it in boiling water, and so I know now that T2 will become 373 degrees Kelvin.

I want to know now what P2 is, and the volume is hardly going to change at all.

It changes because of the expansion coefficient of copper.

And I made a small calculation based on this size, and it turns out that the extra volume that you get because of the heat is only .5 percent of the original volume, so we can forget the fact that V2 is just a hair larger than V1.

It's close enough to say that these are the same.

And of course we have the same number of moles.

So now we can predict with the gas law what the pressure is going to be when we stick it in the boiling water.

So we're going to get that P1 V1 equals n R T1; P2 V2 equals n R T2.

I divide these two equations, lose my n and my R--

we agreed that the volume was the same--

and so we find that P1 divided by P2, T1 divided by T2, or the pressure P2, which is our goal, equals the one atmosphere pressure, which is P1, times T2 divided by T1.

And that is...

T2 is 373, T1 is 273, so it is one atmosphere times 373 divided by 273.

That's what the gas law predicts.

And that ratio is 1.366, so P2 is then 1.366 times one atmosphere.

Now, this gauge is an overpressure gauge, so therefore you're not going to see this number, but you're going to see the difference with one atmosphere.

So what the gauge will show us is 0.366 times one atmosphere, and since it is calibrated in pounds per square inch, it is 0.366 times 15 pounds per square inch, and that is about...

something like 5.4 pounds per square inch.

So that's what I predict, and we'll see how close we get.

You're going to see this, if all works well, on...

Yeah, there it is.

I'm going to set you...

the light situation a little better for you.

I can even make it a little darker.

So, you see here... this gauge.

Note that it is zero, even though it is in ice water.

This valve is open, and so it is at zero because it measures only overpressure.

And now we're going to close this valve.

The valve is now closed.

And we're going to stick it here as this object; that is the volume.

It looks like a bathroom floater to me, in the toilet flush.

That's what it is, probably.

And now it goes into boiling water.

And now look at the pressure.

There it goes--

three, 31/2, four...

That's overpressure in pounds per square inch.

Four and a half.

I'll just give it a little bit more time.

It would take a while, of course, because gas is a very good insulator, so it may take a while for that gas in the copper ball to...

It's already close to five.

We don't have to wait, of course, for the thing to become all the way 5.5, 5.4, but it will probably go up if we wait.

It's now five.

Not much sense in waiting.

What I want you to realize, what happens...

or I want to ask you, actually, what happens if I open the valve now?

Yeah? What happens? Yeah.

So if I open the valve, then one atmosphere, some of the high-pressure stuff gets out, one atmosphere settles inside, and so, since this one is an overpressure gauge, it will be zero.

So right now, it's... it's five.

That's fine with me--

that's very close.

That's within eight percent.

Previous class, we had 5.3, but we were a little bit more patient.

I will open the valve now, and watch what happens.

It goes back to zero.

So some air was let out, because the pressure of the air inside was higher, was higher than one atmosphere.

So, we've seen a test, a reasonable test, first application of the ideal-gas law.

A gas can turn into a liquid, and a liquid can become a solid, and that depends entirely on the kind of substance, the temperature, and the pressure that we have.

And this brings us to the field of what we call phase diagrams.

I will show you your first schematic of a phase diagram, which is also on the Web.

Here you see a phase diagram--

very intuitive.

Here is pressure, here is temperature.

Imagine that we have a cylinder and we put gas in that cylinder and we put a piston on top and we push it down, slowly pushing it down.

So we start here with gas at a particular temperature, which we're not going to change, and we slowly push the piston down.

In this trajectory, the ideal-gas law would hold.

Temperature remains constant, so if you look at the gas law, PV remains constant.

That's called Boyle's Law, by the way, that the product of pressure and volume remains constant.

So the pressure in the gas goes up, the volume goes down, pressure goes up, volume goes down, pressure goes up, until I hit this point.

And now liquid is going to be formed, so the pressure is now high enough at this temperature to form liquid.

If I push further, the pressure will not go up.

All the gas will first turn into liquid, all of it, until the last molecule; and not until everything has become liquid can I push even further onto the liquid, to increase the pressure on the liquid.

It may be a silly thing to do, but I could do that.

You're not going to compress it very much, but you can try.

And in some cases, if you put a tremendous pressure on it, you may turn the liquid into a solid.

And then you reach this domain, where you have a solid.

What is much less intuitive--

that if you did it at a lower temperature and you squeezed the volume so that the pressure would go up, that you may now reach the point here in this phase diagram whereby no liquid is formed--

no condensation of liquid--

but you get immediately the formation of crystals.

So you go from the gas phase immediately into the solid phase.

If you push further down the piston, the pressure will not go up until all the gas has become solid, and then it will continue to go further.

Suppose this were one atmosphere here, and I took some ice--

or you can take a piece of iron at one atmosphere--

and it's a very low temperature, it's a solid.

And I start heating it up, but I keep the pressure one atmosphere.

It's a solid, it's still a solid.

At this point, it begins to melt.

This will be the melting point.

And when I keep heating it, the temperature will not go up until all the solid has been melted into liquid.

Then I can increase the temperature.

Then the liquid will get hotter until you reach this line.

And when you reach this line, some of the liquid will turn into gas.

It will boil at one atmosphere.

You will see it boil.

You cannot increase the temperature.

If it is water, it will stay at 100 degrees centigrade.

There's nothing you can do until all the liquid has become gas.

We call that water vapor.

Then after that, the temperature can be further increased.

So this point would be a melting point for ice and water, and this would be the boiling point at one atmosphere.

So that's the idea behind a phase diagram, and we are going to use them today for some of our experiments.

I have here a fire extinguisher, and a fire extinguisher is filled with CO2--

that's a given.

And I ask myself the question--

seriously, this is really a question that I ask myself; it's not something I made up for you--

I ask myself the question: Is there liquid inside, liquid carbon dioxide, or is there gas inside? And if so, what could the pressure be?

So I measured the volume of that tank.

It's about 40 centimeters high and it has a diameter of about 15 centimeters.

So it's a cylinder.

This is about 40 centimeters and this is about 15 centimeters.

So the volume--

easy to calculate: 2.3 times ten to the minus three cubic meters.

It's clear that it is at room temperature--

that's nonnegotiable--

so that's about 293 degrees Kelvin, the same temperature that we are at.

So now I want to know how many moles I have, the little n.

And what I read on the label that this tank, when it is full, weighs 31 pounds, but when it's empty, no CO2 inside, it weighs 21 pounds...

So I have the mass of the gas--

of the CO2 gas, whatever it is, maybe it's liquid--

is ten pounds.

That's a given.

And that is 4,500 grams--

a pound is 450 grams--

so I know what n is, because the atomic mass number of CO2 is 12 plus 32 is 44.

And so that is 4,500 divided by 44; that's close enough to 100.

So I have 100 moles.

So now I can ask myself, what is the pressure? If this were a gas, what would be the pressure? Well, if this were a gas, then the pressure P would be n times R times T divided by the volume.

I stuck in the numbers, and out comes ten to the eighth.

Ten to the eighth pascal--

an enormous number.

Ten to the eighth pascal is about 1,000 atmospheres, so I doubted very much whether there is gas inside, because I said to myself, at that high pressure, CO2 probably becomes a liquid.

And so I looked up on the Web--

in fact, Dave Pooley did that for me, my graduate student--

looked on the Web, and we found the phase diagram for carbon dioxide.

And what do you see? This is pressure in atmospheres.

It's a strange scale, because it goes five, ten, 15, then there is an interruption and it goes to 73.

And this is the temperature, zero degrees and 20 degrees.

And so at 20 degrees...

Since we know, if there is liquid inside there...

If it were a liquid, then the liquid has to be in equilibrium with the gas.

So you go up here, and you see there's no way the 1,000 atmospheres, the 1,000 atmospheres somewhere there in the corridor.

So already at a temperature...

Already at a pressure of something like 60 atmospheres, it begins to be liquid.

You can't see that 60 here on this scale because it jumps there.

But I called the fire department, and they said it's about 900 pounds per square inch, which is 60 atmospheres.

And so this... this canister contains, then, liquid and gas and stays exactly at that line.

It cannot be higher, it cannot be lower--

think about that--

because gas and liquid at that line exist in coexistence.

And the only way they can do that is exactly at that pressure if the temperature is 20 degrees.

And when you release, when you open the valve, then the liquid CO_2 will turn into gas, but the pressure will always remain 60 atmospheres until you have used up all the liquid, and not until then will the pressure come below 60 atmospheres.

We dealt earlier with the hydrostatic equilibrium, hydrostatic equilibrium of fluids in general, but we used it only for liquids, to calculate hydrostatic pressure.

And the equation for hydrostatic equilibrium--

you will see that, undoubtedly, on the final--

equals minus rho times g.

And this is very easy to use for a liquid, because a liquid is incompressible, so rho is not a function of pressure.

So you can integrate this out very easily, as we did.

You get a linear relation between P and y.

So if I have here y and I have here hydrostatic pressure and let this be y zero--

that is, the sea level, I call that zero--

and let this be minus 4,000 meters, 4,000 meters lower, then the pressure just goes like this.

Here it will be around 400 atmospheres, and it drops linearly.

When you go up to the surface, it drops linearly.

Well, it may be one atmosphere here because that's the barometric pressure, but that's a detail.

I really want this to be the hydrostatic pressure, so this is rho...

minus rho times g times y.

For gases, this would be very different, though, because with a gas, the density does depend on pressure.

And now I will calculate, I will derive for you how the pressure changes with altitude in our atmosphere, and it's going to be very different from this.

And I will do that under the following assumption--

which is not an ideal assumption, but it's not very bad--

namely, that the temperature in our atmosphere is roughly constant everywhere.

And we'll take zero degrees centigrade.

Here it's a little warmer; when you go up, it's a little colder.

We call that an isothermal atmosphere.

What is the density of a gas? Well, it is the mass of a gas divided by its volume.

I take a certain volume and I have N molecules in there.

And each molecule has mass m and this is the volume, so this is the density.

But now I go to my gas law there, and I say, "Aha! Capital N, "which is the number of molecules divided by the volume, is also P divided by kT." So this is P divided by kT times m.

So now I take this equation, and I say, "Aha! dP/dy equals minus Pm divided by kT times g." I bring the P under here and I bring the dy there, so we get dP divided by P equals minus mg divided by kT--

which is some kind of a constant--

times dy.

Let's first talk about that constant.

That constant must have a dimension, one divided by meters, because this is dimensionless--

pressure divided by pressure has no dimension.

dy has the dimension of length, so this must have the dimension of one over length.

In fact, I can calculate what kT over mg is.

That should, then, have a dimension of length.

I know k; I know T; I take zero
degrees centigrade, so T is 273; I know what g is.

What do I take for the molecule... a molecule of air? What is an air molecule? Well, we have 20% oxygen, we have 80% nitrogen.

The atomic mass number of oxygen is 32, nitrogen is 28.

But really, there is more nitrogen than oxygen.

So take 29 as a reasonable atomic mass number for a mean mass of an air molecule, and so you will get, then, that it is roughly 29 times the 1.66 times ten to the minus 27 kilograms.

And you stick that in that equation, you can't be too far off.

And what you find, that this is 8,000 meters.

It has the unit length, or it is eight kilometers, and we call this H zero.

So I will rewrite this a little.

We're almost done with our integration.

So I will rewrite the equation and introduce for that constant one over H zero, because, remember, I turned it upside down there to get length.

So we have dP over P equals minus one over H zero times dy.

I integrate this between P zero, which is sea level, and P at some altitude h, and so dy between zero, sea level, and altitude h.

And that's an easy integral, so I get ln P at altitude h divided by P zero equals minus h divided by H zero, because an integral of dy from zero to h is simply h.

And so what do I find now? That the pressure at altitude h equals P zero times e to the minus h divided by H zero.

This is the altitude in the atmosphere, and if you take this H, then this is the altitude in kilometers.

H zero would then be eight kilometers.

And so if you use this equation, you can calculate what the pressure is at the various altitudes in our atmosphere, and that's not a bad approximation.

Everest is 8.9 kilometers high.

If you use this equation, you will find that the atmospheric pressure is there only one-third of what we have here.

Not enough oxygen to live.

I did quite a bit of observing at an optical observatory in Chile which was at an altitude of 2,400 meters.

At 2,400 meters, the pressure is 3/4 of an atmosphere according to this equation, and water doesn't boil there at 100 degrees centigrade.

Here at sea level, it does, but at 3/4 of an atmosphere, it boils at 92 degrees centigrade, so you can never time to get a soft-boiled egg.

You can never time that properly, because you're used to the 100 degrees centigrade.

In fact, in the kitchen, there were tables which indicated how long you would have to boil potatoes to get them to what you want to at 90 degrees centigrade.

That's all you can get--

you can't go any higher.

On Mount Everest, water will boil at 72 degrees centigrade, so there's no way you can get yourself there some real hot food.

So you need a pressure cooker there, of course.

If you go to 30 kilometers altitude, and you ask this equation what the pressure is, it's 1/45 of an atmosphere, it's only 17 millimeters mercury.

Water at 20 degrees centigrade would boil at that altitude.

I want to show you the phase diagram of water.

This is the phase diagram of water.

And what I want to do is, I want to take some water and bring it to an altitude of about 30 kilometers.

At 30 kilometers, the water at 20 degrees centigrade should start to boil.

And how do I know that? This is the phase diagram for water.

Zero degrees centigrade, 100 degrees centigrade.

This is pressure in millimeters mercury, and the scales are not very clear.

This is all we have.

This is again what Dave Pooley got from the Web for me.

Now, if you take 20 degrees somewhere here, and you have water at one atmospheres and 20 degrees--

which is what we have in this room--

and I'm going to lower the pressure on it...

So I'm going to put it in the bell jar there, and I'm going to take all the air out so that the pressure goes down and down and down, but the temperature is not changing, you go down in the liquid phase, stays liquid... stays liquid...

stays liquid... stays liquid until you reach this point, and then you have coexistence between the gas and the liquid--

we call that vapor, in the case of water--

and that means it will start to boil.

And that happens--

and I looked that up--

at a pressure of about 17 millimeters mercury, which is equivalent to 30 kilometers altitude.

And so we have here some water, room temperature.

Put it here, put it in a bell jar.

This is a 19th-century vintage of a bell jar.

It's very slow.

It will take at least five minutes before we reach that low pressure.

But you will see here the... the wineglass.

There it is.

And we will keep an eye on it, on and off, and when it starts to boil, I can even read the pressure here, but that's not so important.

I want you to appreciate the fact that if you keep pumping long enough that you will hit the line where gas and liquid are in coexistence with each other, and that is our definition of boiling.

So, let me start the pumping and in the meantime, we will do something else, because it will take quite some time before that vintage pump reaches a decent low pressure, which is what we want.

We have to go down to about 15, 20 millimeters mercury.

In the meantime, we'll work on something that is quite similar.

I have a paint can--

you see the paint can there.

You've seen that paint can before.

You remember it--

we evacuated it and it imploded, and we understood why it imploded.

We filled it with one atmosphere air.

This was 25 centimeters, and this was 15 centimeters.

And we took the air out, and then you get an overpressure of one atmosphere, which is one kilogram per square centimeter.

And this front cover alone is about 375 square centimeters, and so the force is huge! And it imploded.

You've seen it happen in front of your eyes when I pumped it out.

Today I'm going to do something more subtle but with the same effect.

I'm going to put in here a little bit of water.

Here is a little bit of water--

there it goes.

And I'm going to bring this water to a boil.

At this moment, there is one atmosphere air in that can.

But as the water starts to boil, the vapor pressure of the water at 100 degrees centigrade becomes one atmosphere.

And so the can fills exclusively with water vapor and drives out all the air.

So the air is gone--

the moment that we see steam coming out, the air is gone.

Then I will tighten it, I will close it, and I will put the can here and let it cool.

What would be the water vapor pressure at 20 degrees centigrade? I just told you that shortly.

That is about 17 millimeters mercury.

That is 1/45 of an atmosphere.

In other words, if this can kept its volume and didn't implode, by the time it reaches 20 degrees centigrade, the pressure in here would be only a few percent of one atmospheric pressure, so it's like having a vacuum in there.

And so clearly, the can will implode.

So we'll try to get the water out, to get it to boil.

I think it is boiling.

I have to make sure that all the air is out.

I really want pure water vapor in there.

Yeah, looks good.

Looks fine.

There it goes.

So now, the vapor pressure goes down, the gas--

if you want to call it the gas, which is what it is--

condenses into liquid, because at lower temperature, it will start to condense.

And what it does, it's going to walk down this line.

Ah! The water is boiling! You see that? Water is boiling.

20 degrees centigrade.

Water is boiling--

okay, so we've seen that.

Let's go back to the...

Shall we go back to water? We started with boiling water here at one atmosphere, 100 degrees centigrade.

As the temperature goes down, it must stay on this line, because water and liquid...

liquid and the vapor are in thermal equilibrium with each other, and as it comes down this line, you see the pressure goes down and down and down.

And by the time that it is 20 degrees centigrade, we would be back at that 17 millimeters mercury.

Now, if this can is leaking, which I think it is, because it should already have collapsed...

If the can is leaking, of course, then that would be a different story.

Then it will not do what we want it to do.

So I may have to try this again.

So, we do it again with another can.

See, that already should have imploded.

So we'll boil this one.

Be a little patient, and we will try it again.

Yeah.

In the meantime, I want to challenge you a little bit and expose you to a demonstration which is kind of bizarre, and I want you to tell me why it doesn't behave the way I want it to behave.

I have balloons, small balloons here, which are filled with air, and I'm going to put them in liquid nitrogen.

So they have a certain volume.

I don't care what that volume is.

They are at room temperature, which is 293 degrees, and the pressure inside is very close to one atmosphere.

There's almost no overpressure in those balloons.

I'm going to make the temperature 77 degrees Kelvin, which is liquid nitrogen.

The pressure will remain very closely one atmosphere.

I think balloons don't have very much overpressure, no matter what you do with them.

And so I want to know what the volume is, how much they're going to shrink.

Well, I apply the ideal-gas law and the number of molecules is not going to change, R is not going to change, so the new volume is going to be the old volume times the new temperature divided by the old temperature.

And so that is the old volume times 77 divided by 293, if the ideal-gas law holds.

And that is one-quarter of V1.

So the volume becomes four times less.

If the volume is four times less, then the radius becomes 60% of what it originally was, because R cubed goes with the volume.

So R2 is about 60% of R1, so it should be very noticeable.

A balloon this big should become this big.

What you will see, however, is something very different.

Okay, I think we are okay on this one now.

We'll try to put the cap back on, if I can find the cap.

Oh, yeah, there it is.

Now, let me tighten it a little better than I did the first time.

And let's see now if it cools...

whether it does better.

Clearly, the other one was never properly sealed.

So, here we have the balloons, and I'm going to dip them in liquid nitrogen.

Come on, can.

It would be quite a coincidence if that one is also leaking.

Normally, they collapse in seconds.

This one doesn't want to collapse.

[can clanks]

There it goes.
There it goes.

It's making obscene noises--

there it goes, there it goes.

Good!

[can clanking]

You will see that--

look at the balloons.

Whew! Here's a balloon.

And here is a balloon.

Anyone's birthday today here? Must be someone.

200 kids--

not kids, 200 grown-ups.

Someone must be... birthday.

50% chance--

no one's birthday? Hard to believe.

Okay, so I predict that if I put it in liquid nitrogen, which I have here, that the radius will become 60% of what it was.

So it shrinks a little--

there we go.

And what you see is something very, very different, and that I want you to explain.

And you have all the tools available.

Keep in mind, I put it in liquid nitrogen--

remember that when you're looking for a solution.

Okay, there's almost no volume left.

It's like a flat pancake--

it's nothing.

Now it comes up, of course, because now it goes back to room temperature.

[can clanks, clatters onto floor]

My goodness, it's having a hard time there.

Why is it not...

[balloon pops]

[class laughs]

Why is it not one-quarter? Why is it so much smaller than what you expect? I'll do one more.

If you come here--

why don't you come here?--

you will see that it is nothing.

The volume is effectively zero.

You see that? Don't be worried.

[class laughs]

And now it comes up.

Last question for you to think about this weekend.

I have here a can with tennis balls, and when you open this can, as everyone knows who plays tennis, you hear...

[makes whooshing sound]

When you go to Europe and you buy coffee, and you open the coffee can, you hear...

[makes whooshing sound]

People like that; they think that's good.

A little bit of vacuum in there, or something like that, makes the coffee stay longer, better, fresher.

Baloney, but it doesn't matter.

In any case, these tennis balls, in the same tradition, you open it up and you hear...

[makes whooshing sound]

Now comes the question for you: Is the pressure inside the can higher than one atmosphere? Or is it lower? It cannot be the same, because then you wouldn't hear...

[makes whooshing sound]

And if so, why would the pressure inside be different? And I'll give you one clue, and the clue is crucial.

You open the can, you don't play with the balls, you wait two days, and the balls are useless.

You can't play with them anymore.

That should give you a clue.

Think about it--

think about the liquid nitrogen balloons.

Have a good weekend--

see you Monday.