Now, today is lecture 2.

And we're going to be talking about counting.

And we're going to be talking about the mul which is a way to count to help us count.

And then-- and this is really exciting-- we're going to be building this.

And that is super exciting.

Now, everybody should have gotten one of these on the first day when we talk about this later in the lecture.

If you don't have this, and for that matter-- oh, there's Laura.

Laura has like thousands of these.

And if you need one, please ask Laura.

And also if you need a goodie bag, which we gave out on the first day, please get in touch with Laura.

And she'll give you one.

We have extras here.

You can stop by our office later, get one please today.

All right, now, last lecture, we left off with-- remember, those guys like Boyle and Priestley and Lavoisier, conservation of mass, right?

And what these guys were doing is they were smashing and burning and pushing on things and trying to figure out what it was that was that sort of fundamental object that you couldn't break apart anymore.

And remember, it was Democritus that said there is such a thing, right?

He said there was a such a thing called atoms, "atomos".

And so we come to Dalton, who said, look, these things we're trying, we're calling them elements right now.

Let's, in honor of that indivisibility idea, let's call them atoms.

And Dalton did a lot of really important experiments.

And he was also trying to smash and break and measure and figure out what are these indivisible-- and what are these things that you can't break apart any more?

What are these atoms?

How many of them are there?

What do they do?

And he came up with a pretty important observation which was that he called it the law of multiple proportions.

And in this, so what Dalton did is he said, look, I'm going to take something and react it with something else.

And it's going to react fully.

OK, so there's no limiting reagent.

There's no excess.

It just everything reacted here and everything and want to measure how much stuff reacts.

And then how much I get.

And so if you take like an example of this law of multiple proportions, if you take an example of how that came about, what he observed is that if you had sort of, for example, maybe you had sort of case one.

And you react some amount.

So he's going to react some amount.

So let's say he reacted-- OK, I don't know, some amount.

Let's write it.

Yeah, some amount of something, maybe, carbon.

And some amount of something else.

How about oxygen?

And he would get some amount, right?

OK, we know he's doing all these measurements.

I'm trying to get you back in that mood.

How much do I have?

I'm going to weigh it, measure it.

That's what he's doing over and over.

But he had some amount of something.

Aha.

Now, he would do this over and over again.

OK, I'm just taking these two elements as an example.

He did this for many, many elements, right?

So some amount of carbon with salmonella.

And in case one, OK, everything reacts.

And I get this carbon to oxygen mass ratio of let's say 1 to 1.33.

I measure it.

That's how much of the sum amount, right?

1 to 1.33 reacts fully.

Remember, no excess, because they add a little more, and it's excess.

No, this reacts fully.

And then you do it, again, case 2.

And he'd find that this same ratio was 1 to 2.66.

Ha, ha.

And he noticed a pattern.

And he noticed a pattern.

And he called it the law multiple portions, because what he found is that you could keep on going.

And in some cases, not all cases.

But in some cases, like carbon and oxygen, for example, these were common ratios, these series.

2.66 divided by 1.33 is an integer.

These things would happen in integer ratios, integer amount.

So I could keep going, and maybe I'd get three or four.

Right?

And what that means is that well, over here, this sum amount of something is it said something about the something, right, because it says maybe these somethings are like CO and CO2, right, for example.

Maybe, aha, but not definitely.

And that frustrated Dalton.

That frustrated him.

We'll come back to that in a sec.

Right?

But he did this.

And he observed these ratios, these integer ratios of what would react with what-- he's trying to figure out what is he making?

All right, what is he making?

Now, he also, as with Lavoisier, remember, I showed you Lavoisier's tableau.

Well, he had his own table.

And here it is.

So here in 1827, he had identified 36 elements.

Now, the thing is, though, OK, 0.1, if you're going to try to brand something, which he clearly tried with these symbols, you can't make them this complicated, right?

I'm pretty happy that that's not how we write our elements.

You can see that this would be very hard.

But he did come up with these symbols.

And you can see it would be very hard to keep going without getting [inaudible].

So I think we're pretty lucky that [? brazilious ?] came along and figured out how to name these elements in a consistent way.

But that wasn't his real problem.

His real problem was what I just alluded to.

See, he knew that he had these different elements.

But he didn't know what the compositions were.

He was so close.

He was so close.

He could see that these-- ah, but is it CO and $CO_2$?

Or is it $CO_2$ and $CO_4$?

That he didn't know.

That he didn't know, because he didn't have an accurate way to measure the mass of an atom, not at

that time.

But he did have this nice ability to see these series.

OK.

But in comes other measurements.

And you had Gay-Lussac who was looking at volumes, right?

And so Gay-Lussac was saying, well, look, if I take a certain amount of volume of a gas, one gas, and I reacted again completely, no access, OK?

And I reacted with another gas.

And then I see I get something else over here.

I can measure the volumes.

And I do this for the cases where there's nothing left over, nothing limited it.

And so he would find, for example, what's shown up there.

He would find that you've got for every two volumes of hydrogen gas-- so you've got two volumes of hydrogen that he would react with one volume of oxygen.

And he would notice that he would get two volumes of something that had-- let's not put a dash there-- it had hydrogen and oxygen in it.

He knew that.

Now, he knew this because he was very careful.

So nothing leaked.

He said no.

Everything is accounted for.

Lavoisier holds-- conservation of mass.

I'm not losing anything here.

And as Lavoisier said, in a chemical reaction, mass is neither created nor destroyed.

So I know that whatever is over here must consist of whatever I have here-- oxygen and hydrogen. And he kept doing it for other lots of different types of gases.

But it was very strange, because what happens is sometimes, the volumes are the same.

Like he did it for hydrogen and chlorine.

One volume of hydrogen, and one volume of chlorine would give you two volumes of hydrogen and chlorine-- something that had hydrogen and chlorine in it.

And so there, you'd mix one and add another one.

And you'd get two.

It seems it would kind of makes sense.

You add one volume.

But here, you'd mix two volumes, add another volume.

And it would feel like it was reduced.

Something in the reaction happened that gave you a lower volume.

What could it possibly be?

Right, what was going on.

This is what they were grappling with.

This is where they were.

And then along comes [sighs] Avogadro.

Now, I wish we all called him Lorenzo Romano Amedeo Carlo Avogadro which is his full name.

But we shall call him Avogadro.

And he made a critical proposal.

He said, look, the way to explain what's going on in Gay-Lussac's volume measurements is to make the following assumption-- oh, I have another board over here-- is to make the following assumption.

And I think if you've had high school chemistry, you know what this is.

But I'm going to put it here, because of how important it is.

So Avogadro said that at constant temperature and pressure equal volume equals same number of particles.

So if I have the [? efomethesine ?] temperature [inaudible],, if I have the same volume, I have the same number of [? particles. ?]

Now, what are those particles?

Oh, he was flexible.

This is important.

He was flexible.

He said, those particles could be atoms, or they could be combinations of atoms.

They could be molecules.

Remember, at the time, they hadn't all settled on the atom as a name, even.

What were these things?

Equal volume, same number of particles.

And this really helped, because you see, if you make that assumption, and you believe these volume measurements, then it becomes very clear what you have there.

Right?

If I have the same number of particles in here, and I believe in Lavoisier, who tells me nothing comes in or out in terms of the math, I didn't create or destroy atoms, then you know that if I have two volumes of water, let's say I started with o and h.

It doesn't work.

I can't make enough water.

I just can't make enough particles.

The number of particles on the right has to be equal to the number of particles in those two volumes of hydrogen. It's got to add up.

Counting has to work.

That's what Avogadro gave us is this idea of counting.

And that's why we've named a constant after him.

So I'll get to-- Gesundheit or in honor of Avogadro, "salu-tay".

[laughter]

He was Italian.

And so if you believe Avogadro, if you believe Avogadro, this becomes clear, because the only way for these things to work then is if here I'm counting how much volume's here.

So here I got HCL only if this is a two, and that's a two-- diatomic gases.

That's the only way the numbers work.

And here, I've got to have two and two.

And this must be H2O.

Oh, or at least it's one simple and very plausible explanation if Avogadro's hypothesis was right.

Dalton didn't buy it.

Dalton said, uh, uh, no way.

You've got leeks, dude.

And he really was harsh.

Everybody thought water was OH at the time.

And so Dalton said, you must have made a mistake.

I block you on Instagram.

[laughter]

I block you.

And it was pretty tough, because Avogadro put a lot of really brilliant work out there that was really ignored for 50 years until Cannizzaro came along and others.

50 years later, two years after he died was when we had the first evidence that the scientific community was willing to believe that his hypothesis was right.

What a big deal his work was.

OK.

Now, I mentioned that this is-- I'm mentioning all this because we are going to use his constant a lot, the Avogadro constant.

And it's a way-- and so I want to talk about that constant-- and I want to do it in the context of our candle.

And so because this constant, this number, this idea of the number of particles you have in chemistry is the link between the atomic world and the macroscopic world.

That's what his constant does for us.

OK?

Now, let's put that into context.

If I were to ask you a question like this, I have that candle.

Remember, the candle, the $C_{25}H_{52}$.

And I say, well, OK, you've lit the candle.

How many candle molecules burn in a minute?

I want to know that.

Don't we all?

Right?

That's a pretty important question.

Why are people nodding?

Thank you for that agreement.

And I could say, well, I assume the 12-hour candle of the mass and 90 grams, let's get that off of a candle on Amazon.

Right?

How many molecules burn in a minute?

What I need to know is, well, how many are we sort of talking about?

How do we know how to even answer this question?

And it comes from the link.

But see, to answer this kind of question, which I'm sure so many people wanted to know back then, you must have a way to talk about big numbers.

You must.

Why?

Because there's a lot of these molecules in a very tiny amount of volume-- a lot.

Drop of water for perspective has 10 trillion molecules-- 10 trillion molecules.

Anybody know how many cells you have in your body?

600 trillion?

Let's call it 100 trillion just to be nice.

Right?

It's a 100 trillion cells.

And so if you're going to talk about how many cells you have, or how many molecules are in a drop of water, you need to have a big number, a big scale.

And that is what a mole is.

Avogadro didn't come up with this number.

This was gotten much later, right?

But we name it after him in honor of the work he did on counting atoms and molecules.

And that's the number.

This is the way we talk about it.

And just to be complete, Avogadro's constant, which is a mole, one mole is equal to 6.022 times 10 to the 23rd.

It's just a number.

All right?

It's just a number.

But it's a really important one because it provides the link.

OK, if I take like how many eggs are there in a dozen eggs.

It's 12.

How many bags of sugar are there in a dozen bags of sugar?

It's 12.

All right?

How many atoms are there in a mole of atoms?

That number.

That's all it is.

It's a number.

OK?

Oh, but a dozen eggs weighs differently than a dozen bags of sugar.

So a mole of one atom will have a different amount of mass, a different weight than a mole of another atom.

And that's the link that he gave us.

That's the link that this number gave us.

And, in particular, again, going back to this idea of counting, we can count now with this link.

If I said, how many pennies are there?

And I gave you a mass of pennies, right?

If I had 2.5 kilograms of pennies, and I say how many pennies do I have?

Then all I need to know is how much each penny weighs.

I'm not going to count them.

Right, you don't want to count them.

You want to do something a little easier.

So if I have one penny, weighs 2.5 grams.

Then I know that I can just get this really easily.

I have 1,000 pennies.

OK.

What a mole does-- a mole gives us the link.

So a mole, if I have moles, then I can go to the mass of an atom-- the mass of an atom.

And from that, I can go all the way up to grams.

And this is critical.

OK.

The mass of the atom and the grams per mole are the same number.

That's what the mole gives us.

Right.

So in the case of carbon, carbon has a mass per atom of 12.011 AMU.

These are atomic mass units.

It's just a way to measure things that have very little mass.

But the link here is that if I have a mole of these things, if I have a mole, one mole of carbon, then I have exactly 12.011 grams.

Back and forth, grams, atoms.

Numbers of atoms, mass per atom.

All right, so one carbon atom has a mass of 12 atomic mass units.

We don't need to know.

This comes from weighing neutrons and protons.

All right, we'll get to that next week.

We don't even know that yet.

But this is a mass of things that the very tiny scale.

And if I have a mole of those atoms, I get to grams.

OK?

So it's my link.

Now, where is this amount?

Where is this critical number that gives me the counting per gram?

It's, of course, here in your periodic table.

This is where you get that information.

OK, so we'll be building that up as we go.

Well, let's say why do I need to count?

Well, let's go back to answer our question now.

That's the link.

That number, Avogadro's constant links these atomic masses to grams.

All right?

Now I can answer the question.

How many candle molecules burn in one minute?

Well, let's do that.

Because now I know that one mole of $C_{25}H_{52}$ is something like 352.7 grams.

How do I know that?

Because I can look up what a mole weighs of carbon in here, how many grams.

And I can look up how many grams of mole hydrogen weighs.

And then I can multiply 25 and 52.

And that's how many grams I get.

One mole of it is that.

Aha, the link is coming.

The link is coming.

It's very exciting.

Because in a 12-hour candle, if it's 90 grams, so if I have a 90 gram candle, as I've said in the problem, then I know now that that is about 0.25 moles.

Right, because of this.

See that?

90 over that.

So now, I know how many moles of candle I've got.

I know how many moles of candle I've got.

But remember a mole is just a number.

It's just a number.

So this is how many moles of candle molecules I have.

And OK.

In one minute.

It says in one minute.

So in one minute-- well, OK, there's 60 minutes in an hour, and it's 12 hours, just a 720-minute candle.

Right?

So in one minute, 1/720 of that is what burns, if it lasts 12 hours.

I'm just taking the information I was given.

And I'm using the math of minutes and hours.

1/720, OK.

All right, good.

So 1/720th of the candle burns.

In one minute, 1/720 of the candle burns.

And if the candle is that many moles, that's this many moles-- 0.003.

No, 0347 moles.

Why did I do all that?

Because I can now answer the question, cause of Avogadro's constant, because a mole is just a number.

A mole is just a number.

So if I've got that many moles, then I know that I've got 2 times 10 to the 20th molecules roughly.

So I have been able to answer this question by going back and forth.

[inaudible]

I look here to get the grams per mole.

I figured out how much candle I had.

In grams, that gave me how many moles I have of candle.

And then gave me the number-- back and forth, back and forth.

That's what we do.

Avogadro's constant gets us in and out of the atomic world.

OK?

And so we're going to practice that a little bit.

And this gets me back to what we started on Wednesday.

And I always like to give you guys advice if I can.

And last time I mentioned, if you're on a date at a restaurant, ask them to give you, not a candle, but some $C25H52$ as a test for the quality of the restaurant.

[laughter]

Now, I'm asking you to do something else.

It's noisy.

You want to talk.

Right?

You go out on a first date.

You're a little nervous.

You want a quieter room.

Ask them if there's a room where you can kind of shut the door and get that noise out.

And it's quiet so you can talk.

But now, you are armed with very important knowledge, because now, you know how much volume you have of air in the room and therefore, how much oxygen you have.

[laughter]

And if you light a candle, and you're taken 10 to the 2 times into 20 the molecules of candle out of the year, by the way, how many molecules of oxygen are you burning?

Ah, how many molecules of oxygen?

Ha, for at 38, balance-- 38 times as many molecules.

For each one of these that burns, you need 38 molecules of oxygen. So 38 times as many O2.

Now, I don't know about you, but if I'm on this date, I'm looking it up.

The human-- [laughter]

--needs 1.9 grams of oxygen per minute.

You need that.

I mean, maybe it's variable.

But you need roughly-- and there's two of you.

0.4 grams of O2 is burned per minute.

0 point-- who's going to win-- you or the candle?

So you got to be just aware of this.

And now look, This is kind of thing we can do now.

And I'm not saying that everyone's going to understand this.

You might not be with an MIT student.

You might be with a student from some other school down the street.

[laughter]

All I want to say, be nice, be nice.

It may not be their fault that they're where they are.

[laughter]

It may not be their fault. Go through this with them carefully, slowly.

[laughter]

Explain these concepts.

And this gets me to the other topic of conversation that you can have.

And this gets me to my why this matters moment for today.

Why does this matter?

Why does counting and thinking about how much of stuff you're using by using chemistry and moles?

Why does that matter?

Well, let's take an example.

Let's take an example of nothing less than the number of humans on this planet.

This is the population of humans.

And you can see that it wasn't all that much until lately.

But let's focus in on this part here.

Let's focus in on that part there.

Right.

So that's billions of people.

So if I zoom in on this, there's a change.

Look at that.

It's kind of going a little bit up.

But now, it really kicks up.

What happened during that time?

What happened is a very important chemical reaction, became easy to do or a lot easier.

OK?

And the process that enabled that is called the Haber-Bosch process.

But it is what allowed us to feed billions of people in a sustainable way.

It's arguable whether we're doing that sustainably but in a way where we could actually produce enough-- enough what?

Enough ammonia.

And here's the deal you see, because plants need nitrogen to grow.

Right?

And there's 70% of the atmosphere is nitrogen.

But it's useless, because it's N2.

And N2 is one of the strongest bonds in nature.

And plants can't break it.

Now, we knew how to break it before.

We knew how to do this before, right?

This is called fixing-- fixing nitrogen. And the way it works is you go N2 plus h.

This is a reaction, goes to NH3.

What is wrong with this?

[inaudible]

It's not balanced.

Thank you.

Did anybody say, balancing is the same as mole ratio.

Balancing and mole ratio's the same-- counting.

Remember, we talk about that.

But balancing is counting.

So let's see-- oh, 2 here.

OK, that helps me with the nitrogen.

Oh, but this is H2.

I meant to put that in the first place.

And then so this is a 3.

That reaction was known and doable.

But it took tremendous amounts of energy.

So it's very difficult to scale.

And Haber-Bosch came up with a way using catalysis to do it at much lower temperatures.

Catalysis is something we'll learn later in the semester.

But how do we answer this question?

What's my question?

My question is how long can we keep going?

50% of every protein you put in your body-- 50% come from this-- comes from some plant that was grown using this process.

That's how important it is.

It's 500 million tons of nitrogen is made this way every year.

Well, it's just counting.

One mole of ammonia is 17 grams.

Oh, we're doing the same thing.

How did I know that?

Periodic table.

That's how I knew that.

One mole-- if that many molecules of ammonia, I've got 17 grams of it.

Let's say I just need to make the same amount of-- it says so times 10 to the 6 tons.

I need that much tons of NH3 per year.

That's our-- yeah, per year-- per year.

OK, so that sets up my problem.

I know how much I need.

I know how many atoms we're talking about in one mole.

Now, I can actually understand how many moles-- so I'm just going to not do the detailed math but

how many moles I have.

So I've got 30 times 10 to the 12th moles.

This is how many moles I needed per year.

How much do I have?

Well, we know how much the atmosphere weighs-- it turns out.

We know how much the atmosphere weighs.

The mass of the atmosphere is something like 5 times 10 to the 21st grams.

And if I take 78% of that as my N2, so 78% is N2, then I can tell you that I've got-- let's see, 1.4 times 10 to the 20th moles of N2.

That's how much I have available.

OK?

So if I keep on taking N2 out of the atmosphere, then I can now answer the question just like the candle.

All right, how long do you have?

I can answer the question, how long can we keep taking N2 out of the atmosphere and using Haber-Bosch, right?

Oh, well, you would use the balanced reaction.

So for every mole event do I take, I get two moles of ammonia.

That's good.

And you can work backwards.

And you can learn that you're good.

We're good for now.

We have roughly 100 million years.

OK?

We have 100 million years that we could keep consuming.

And then we'll run out of N2 in the atmosphere.

How fun was that?

OK, I love doing this.

Now, when I get excited-- wow.

Wait, before I get excited, what's the limiting reagent?

[? add ?] enough N2.

Is it N2?

How do you know?

How could you show?

How can you prove?

Limiting range, it means what runs out first?

Well, oh, I'll just take H2 from the atmosphere, right?

Na, uh.

There's no H2 in the atmosphere.

Where's the H2?

Di da, it's over here-- H2O.

It's in the oceans.

How much H2O do I have?

Well-- [laughter]

It turns out you've got a lot.

10 to the 23rd-ish moles of $H_2O$, right?

10 to the 23rd-ish moles.

And so because of those coefficients of the reaction, you know exactly which one is going to run out first.

You know it's $N_2$, so your answer over there was correct.

$N_2$ is the limiting-- I love that little dance.

I'm going to try to learn that.

I'm going to learn that later.

If you try this, you go 100 million years.

You run out of $N_2$-- plenty of $H_2$ left.

Let me read you.

When I get excited, I didn't miss this.

I like to throw stuff like T-shirts.

Now, the thing is these T-shirts are not-- [laughter]

Now, and I go like this.

[laughter]

And then I go all the way-- ahh, to the back, all the way up there, all the way up there.

All right.

OK, now look-- why do I throw T-shirts?

I have no idea.

[laughter]

Sometimes things are just better left unknown.

But please, I'll throw T-shirts 34 more times.

Every lecture, I'm going to throw some T-shirts when I get a little excited.

And please, if you get a T-shirt a second time, share.

Right, don't take it-- give it to somebody-- trade.

Right?

Sharing is caring.

Now, where were we.

OK, now, here we are.

We got to build this?

Here's the thing-- in 1669, there are only 12 elements known.

But yeah, I hope I've given you a sense.

All these people and many more, I haven't even named.

We're working on discovering new ones.

By 1799, we had 34.

Dalton and I mentioned had a table of 36, and they kept going.

And as we got more and more elements, a lot of people were wondering, how do we organize these?

How do we organize these?

And are there patterns?

Are there patterns, right?

That became a hot question.

Yes, they still wanted to smash and break and burn and make new, discover new things.

But they also wanted to know what it all means.

It wasn't just like a few-- it was a lot.

And so they started looking for patterns.

And one of the first people was Dobereiner, who came up with what he called the Law of Triads, which was a law that worked three times.

[laughs] But what he found-- [inaudible],, everybody was arranging these elements in different ways.

[inaudible] arrange it by mass.

No, arrange it by something else.

And he found that if you listed him kind of by mass-ish, then the middle atom of the three would weigh exactly half or roughly half of these two.

That was cool for three times.

OK, Newlands did something really important.

Well, first of all, he loved music.

And he arranged elements like this and again, kind of by mass.

And he noticed that there were repeating patterns every eight.

Now, unlike a lot of triads, this was really important, because the Law of Triads, you said, look, I can take these three.

I can take those three.

That's kind of cool, but it wasn't periodic.

It didn't happen these three then, the next three.

And whereas the Newlands, it was periodic.

And it sort of also only worked for 2 and 1/2 runs.

The Law of Octaves-- he called it the Law of Octaves.

Oh, this was the time to look for patterns.

And many people took a stab at it.

Julius Meyer made some really important contributions in measuring the molar volume by knowing the densities and finding patterns.

they were looking for periodicity.

But as many of you may already know, it was Mendeleev who figured out what needed to be done.

And what he figured out was absolutely critical.

Let's go to the middle.

And notice, he's listing them-- OK, he didn't know whether to put it this way or sideways or what.

This is oriented differently than what we do today.

But it's OK.

He's just listing them and looking for patterns.

But what Mendeleev did that were so critical-- what he did that was so important is that he organized them.

So Mendeleev-- Mendeleev-- he had, at the time, this was-- let's see 1869.

He had 63 elements.

And he called this a periodic system.

And the critical piece is he arranged them by mass.

And this is so important-- properties.

He really believed that the periodicity was the driver, and that the properties were critical in this organization-- the properties and the mass.

Right?

He was so certain that you had to take them both into account that he actually left gaps.

That was new.

At least for this many elements and arranging it this way, that was different.

And so like, for example, look at this.

Over here, so here you have aluminum, silicon.

And you know what else?

Uncover this, so you can it.

Oh, zinc-- OK, zinc.

Right, arsenic.

But he said, no, don't just put this next to zinc because of the properties.

This doesn't have the right properties.

This should be more like phosphorus.

So we're going to leave gaps.

And this is really important-- he predicted the properties of those missing elements.

So if you call this, for example, with silicon, we know this is germanium.

We know that now.

But back then, no one had smashed and burned and broken enough stuff to find germanium.

So he said, well, we can predict it.

Because it comes here in this table in this organization, therefore, it must have properties similar to other elements in that column.

And he called it eco-silicon.

Eco just means one next, right?

And so these are predictions by Mendeleev-- predictions of the properties of this missing element.

And here's what you get when you actually find germanium years and years later.

And you measure those properties.

And look at that-- it's incredible.

Right?

Thank you for that wow because I felt it too.

Any more T-shirts?

Aw, I felt the wow, and it made me excited.

[laughter]

All right.

This is a big deal.

You go back to here-- iodine and tellurium-- no, I'm not going to give you the mass ordering, because it doesn't make sense in terms of property.

So I'm going to keep the mass the other way, even though this has a higher mass, look at that question mark.

I'm still going to put it before because it makes more sense.

That's what Mendeleev did, and it was absolutely critical.

It really brings us to the current system that we use.

So this is now the periodic table.

Our periodic table today is based Mendeleev's organization.

And we're going to be spending a lot of time with it.

So we're not going to go through everything.

But a couple of things.

First of all, there's a key.

That's the key to the table.

If a table doesn't have the key, don't use it.

It should have a key.

And so, for example, oh, look at that number.

That looks nice.

Right?

That's the AMU or the grams per mole which we now know are equivalent-- not equivalent but relatable.

You have a mole, that's how many grams you have.

Or if you have one atom, that's how many AMU's you have, right there.

The key tells you what things are.

Right?

Now, if you have a more complicated one, don't panic.

Here's yours.

This is your periodic table.

Ha, look at this.

It's all in the key.

There's a lot more information in here then in the basic ones.

But it's all in the key.

It's all described in the key.

And the key may be different on the front and the back.

There may be different properties included.

But it's all in the key.

And look at it.

Some of these we know like the weight-- atomic weight-- that's what we've been talking about.

Melting point sounds kind of obvious-- boiling point density.

And then you get to these other things-- electronegativity, first ionization.

All these things, what are they?

We will find out literally next week, all right, as we build up our understanding of the atom.

But it's all on the key.

Now, this is so important that I really ask you with passion to carry it with you everywhere you go.

I really do, because you never know.

Remember the restaurant example.

You never know when you need it.

You never know.

And, in fact, you can send me pictures.

Students do this.

I love it.

You don't have to send me pictures.

You can just post it on your feed or whatever it's called.

[inaudible]

What's it called?

I don't know-- a thing.

So they took it all the way up to Dartmouth, because there were some people up there apparently

that needed help with their chemistry.

[laughter]

Here we are at a restaurant, looking at the periodic table.

Here's another.

It's a good thing to have at a restaurant.

You never know.

Here it is again.

Look at how happy they are.

Look at that, right?

They're really getting into it.

They're really getting into it here, teaching family members about this or being proud of it with the whole family.

[laughter]

Take it with you.

It'll come in handy no matter what.

Here it is under a Christmas tree.

Right?

It's in here somewhere.

And there it is.

And look, even if you're dressing up for a theater, it's going to be important.

OK, now, this is the one I want to get to-- OK, even if you're a little tired.

It's going to be what comforts you.

I think it's being used as a pillow.

That's OK, it's comfortable.

Look at this.

[laughter]

Right?

Right?

You know even at the dance club, even at the dance club, oh, I see that.

You might need this.

You don't know.

You don't know.

Look at how happy he is.

[laughter]

Look at that, because he has this.

Take it with you everywhere you go.

Now, there's a bunch of things about this table that we'll be talking about.

We just built it up.

The first things is just some nomenclature.

Right, just how do we name things?

We name things that the rose, we call them periods, and there's seven of them.

The cones, we call them groups, and there's 18 of them in the way we organize, in the way we present it, right?

But there's something kind of important that happens here which is that these rows down here are actually 14 elements longer.

So right, so this is 18 across.

But these guys fit in right here.

They go right there.

Right?

These are called the lanthanides because that's lanthanum-- innovative naming.

These are called the actinides.

OK.

They go in there in between.

Why don't we just put it all there?

Well, you can.

But then you got the bottom rows are 34.

Right?

And then the next one's up.

There's only two rows in the periodic table that are actually 18.

Then because then you got 18 here.

And then these are eight.

And then you got two up top.

Now, why?

That we will understand when we understand the atom a little better next week.

OK?

But for now, I just want to call your attention that these are the names that we use.

And those top ones are called mean group.

These ones are transition elements.

You have alkalines here.

You have lanthanides, actinides.

These are the names that we give these regions that get more clear when we know what electrons are.

And I think I'll tell you just the last bits about just a couple more observations about the periodic table, I think, when we start our lecture on Monday since it's 11:54.

And as promised, I want to let you guys out right on time every lecture.

Have a fantastic weekend and see you Monday.