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5.111 Principles of Chemical Science, Fall 2008
Transcript – Lecture 2

PROFESSOR: So, our first question here is about limiting reactants. So, that's something you will encounter in your review reading for the sections, that kind of review -- what we hope you have picked up from high school or will pick up quickly by doing some review. So, how about we have everyone take ten more seconds on the clicker question, get your final answer in here.

All right. So, let's see what we have. All right, so it looks like we weren't showing the percentages here, but it looks like hopefully most of you were able to get the correct answer of H2 being the limiting reactant. It looks like we're still figuring out -- this room was just renovated, we're still working out exactly how the electronics work. So, normally we'll see a percentage of how many of you got it, but I'm going to say it was probably about 95% got the answer right. So, good job there. If you didn't get the answer right -- we'll send these questions to your TA, so any time you get a clicker question wrong and you're confused, bring it up in the next recitation section and you'll be able to discuss it there.

So, starting in, we can switch over to the to the notes now. When we left off on Wednesday, what we had really been doing is trying to give you an overview of all of the different topics that we're going to be going over this semester. And also, to make a couple of those connections between the principles we're learning, and some of the exciting research that's going on at MIT in the Chemistry Department, and also, to give you the idea that we are going to be trying to make these connections between the chemistry and things like human health or medicine.

So, now we get to actually take a step back and start at the beginning, because before we can talk about some of the more complex issues, which involve interactions between molecules reacting or even when we're talking about individual molecules -- the bonds that form between individual atoms -- before any of that we actually need to establish a way that we're going to describe and think about how an individual atom behaves. And the way that we'll do this is starting with talking about the discovery of the electron and the nucleus here.

Once we go through that, we will be able to talk about describing an atom using classical physics. So, once we have an atom and a nucleus, what we'll try to do is apply the classical mechanics to explain how that behaves. What we'll find is that this fails, and once this fails we're going to need another option. Luckily for us, we have quantum mechanics, which we'll be talking about for the next few lectures, and we'll dive into that. We might get a chance to introduce it today, but certainly in next class we'll be introducing this new kind of mechanics that's going to allow to describe the behavior of atoms.

So, I want to point out that it makes a lot of sense for us to start with the discovery of the electron and the nucleus, because it really highlights one of the big issues that comes up in all chemistry research that you do, and that is how do we actually study, or in this case, how do we discover atoms or sub-particles that we actually can't see at all. And there are lots of solutions that chemists come up with -- there's always new techniques that allow us to do this, and these are just some of the first, and we'll go through them in a little bit of detail here.

So, this all starts, in terms of putting it in its historical context at the turn of the Century, we said we'd start right in on the 20th Century of where chemistry was. And where we where at the start of the 20th Century in the late 1890's is that we were at a place where there was great confidence in our understanding of the universe, and our understanding of how all matter worked.

So, people in the chemistry community and in the physics community had this general feeling that the theoretical structure of the entire universe was pretty well understood. And they had this feeling because there had just been this huge boon of discovery, of scientific advances that included Newtonian mechanics, it included Dalton's atomic theory of matter, also thermodynamics and classical electromagnetism.

So, you can understand they really felt quite confident at this time that we could explain everything that was going on, and in fact, a really telling quote from the time was said by a professor at the University of Chicago, and what he said is, "Our future discoveries must be looked for in the sixth decimal place."

So, basically what he's saying here is we pretty much understand what's going on, there's nothing new to really discover, all we need to do is measure things more precisely. So, that's not exactly the case, and we're going to start in at the point where right around this time of great confidence of feeling all has been conquered, there are some observations and discoveries that are made that completely break down these theories.

For example, in terms of the atomic theory of matter, at the time at the turn of the Century, the understanding was that atoms were the most basic constituent of matter, meaning you couldn't break atoms up into anything smaller -- that was it, you're done. And with using Newtonian mechanics, it was assumed since this type of mechanics worked so well to describe everything we could see, it could even describe the universe and planets, that, of course, we could use Newtonian mechanics to describe how an electron -- actually, we didn't even know about an electron here, but how atoms behaved, and it turns out this is not the case, and the first step in discovering this is not the case, was accomplished by J.J. Thomson, and J.J. Thomson is credited for discovering the electron. He was a physicist in England, and what his laboratory was studying is something called cathode rays, and cathode rays are simply rays that are emitted when you have a high voltage difference between two electrodes.

So, if you look at this set up, what he did when he was studying these rays is he had an evacuated tube, which is schematically shown here, where it's evacuated of all it's air and filled instead just with hydrogen gas, and he had this high voltage difference between an anode and a cathode, and he actually put a little hole in the anode here, so these cathode rays that were produced could shoot out of the cathode and actually could be detected as this luminescent spot on a detector screen.

So, lots of people were studying cathode rays at the time -- one reason is they actually gave off this bright glow -- if you put them in an evacuated glass tube, you got these crazy patterns and glowing colors. So, for that reason it was a very hot issue in terms of research. But also, no one really knew what these were and Thomson was seeking to figure out some more properties of them, and he had the theory that maybe they were actually charged particles of some sort, and others had proposed this in the past, but they didn't really have an experimental set up to test it. And that's what Thomson did. And what we did was he put two detection plates on either side of these cathode rays, and when he put a voltage difference between these two plates, he wanted to see if he could actually bend the rays and test if they're actually charged or not.

So, when the voltage difference between the plates is zero, or when we just don't have the plates there at all, the cathode rays are not bent, they just go right in a straight line, and they can be detected on this screen.

When he actually cranked up the voltage between these two plates, what he saw was really amazing to him, which is that he actually was able to bend these rays -- this had never been observed before in any capacity, and he was able to detect on his screen that there was this deflection, and he could even measure the degree of the deflection that he had.

So, we know now that we have charged particles. Are these negatively or positively charged, based on this evidence?

STUDENT: Negatively.

PROFESSOR: Yeah, that's right. So, what we have here, cathode rays we now know are negatively charged particles. And, in fact, he named these negatively charged particles. Does anyone know what he named them? No, not electrons -- very good guess. He named them corpuscles. Has anyone heard of corpuscles? A little bit. Yeah, so it was later named that these particles were, in fact, electrons, and that's what they are. J.J. Thomson continued to call them corpuscles for many, many, many years after everyone else called them electrons, but I'm sure no one minded because he did, in fact, discover them. And he was actually able to find out more than just that these were charged. From classical electromagnetism, he could actually relate the degree of deflection that he saw to the charge and the mass of the particles.

So, using that he could say that delta x, and we'll put sub-negative, because we know now that these are negative particles, is proportional to the charge on that particle over m, which is the mass. So, we have e being equal to the charge of the negative particles, and m, of course, is equal to the mass of those particles.

So, Thomson didn't stop here, he actually continued experimenting with different voltages. And what he found was if he really, really ramped the voltage up between those two plates, he could actually detect something else. And what he could detect here is that there was this little spot of luminescence that he could see on the screen that was barely deflected at all -- certainly in comparison to how strongly this first particle was deflected. The second particle was deflected almost not at all. But what he could tell from the fact that there was a second particle at all, and the fact that it was in this direction, is that in addition to his negative particle, he also, of course, had a positive particle that was within this stream of rays that were coming out.

So, of course, he can use the same relationship for the positive particle, so delta x now of the positive is proportional to the charge on the positive particle all over the mass of the positive particle.

So, this is interesting for several reasons. What did he manage to pull out information-wise from using these two relationships? And actually to do this, he made a few more observations. The first, which I just stated, is that the deflection of that negative particle was just far and away more extreme, much, much larger than that of the positive particle. The other assumption that he made here is that the charge on the two particles was equal.

So, how could he know that the charge on the two particles was equal? And actually he couldn't exactly know it -- it was a very good assumption that he made, and he could make the assumption because he, in fact, did know that what he started with was this hydrogen gas. So, he was starting with hydrogen. If some negative particle was popping out from the hydrogen, then what he must be left with is h-plus, and since hydrogen itself is neutral, the h-plus and the electron had to add up to be a neutral charge. So, that means the charges of the two pieces, the positive and negative particle, must be equal in terms of absolute charge.

So, using this relationship, he could then actually figure out by knowing, which he knows how much each of them were deflected, he could now try to think about whether or not he could make some relationship between the masses -- between the mass of the positive and the negative particle.

So, this relationship he was looking at was starting with the deflection, and the absolute distance that the particles were deflected. So, what he could set that equal to is he knows what x is proportional to in terms of the negative particle, so that's just the absolute value of the charge over the mass of the negative particle. He could divide all of that by the absolute value of the charge of the positive particle, all over the mass of the positive particle. And as we said, he made the assumption that those two charges were equal, so we can go ahead and cross those right out. So, what that told him was if he knew the relationship between how far they were each displaced, he could also know something about the relationship of the two masses. So essentially, there was an inversely proportional relationship between how far the particles were displaced, and what the mass of the two particles turned out to be.

So, because he, of course, observed that the negative particle travelled -- it was deflected much, much further by those plates, what he could also assume and make the conclusion of is that the mass of that negative particle is actually larger or smaller?

STUDENT: Smaller.

PROFESSOR: Much, much smaller, exactly, then the mass of the positive particle.

So essentially, what he found here is the relationship between the mass of an electron and the mass of the rest of the atom, the rest of the hydrogen atom there, which is an ion in this case. And, in fact, it's so much smaller, it's close to 2,000 times smaller, that we can make the assumption that essentially the electrons take up no mass. I mean they take up a teeny bit, but essentially, when we're thinking about the set up of the atom, we don't have to account for them as using up a lot of the mass we're discussing.

So, Thomson came up with a model for the atom due to this, and this is called the Plum Pudding model of the atom, and he was, as we said, English, so plum pudding is kind of a British food. Has anyone here ever had plum pudding? A couple of people. Okay. I've never even seen it, so that's good -- you must be better travelled than I.

So, the idea that he had here was he treated the whole of the atom as sort of this positive pudding, so the majority of the atom was just kind of this goopy, positive stuff that you could think about, and within the pudding, he had all these negative charges, which were the electrons, and they were the raisins or the plums that were in the pudding.

So this was a revolutionary model of an atom when we thought of the fact that before this experiment, the understanding was an atom could not be divisible into smaller parts, and now here we are with subatomic particles with electrons, and this wonderful Plum Pudding model. So, for those of you that haven't actually had plum pudding, which myself is included, I threw a picture up here. This was my first glance at plum pudding, and I guess you can see that this must be that positive part -- most of the plums are within that, and you can see all these little raisins or plums in here, that would be that negative charge.

So, that already was a big advancement from where the understanding was at the time. We already moved way forward and completely revolutionizing the understanding of an atom in that there's something in an atom -- it's not the smallest thing there is.

However, as you know, we didn't stop at the plum pudding model, which is good, because it's a little goofy, so it's nice to move on from that and move on we did. About 10 to 15 years later, another physicist, Ernest Rutherford, actually put this plum pudding model to test, and he did it through studies that he'd been doing on radiation that was emitting something called alpha particles.

So, Rutherford, some of you may recognize that name, is a very famous physicist who made a lot of contributions in terms of radioactivity. When he was studying these alpha particles, he was actually the first person to identify the difference between different types of particles that radioactive materials emit. And he got this particular material that he was studying, radium bromide from his good friend, Marie Curie, who, obviously, also was a leader, really the leader in figuring out much of how radioactive materials work. She has two Nobel Prizes for her work in radioactive materials. And something that maybe many of you think, which I know I always think when I hear about radioactivity studies in the early 1900's, is oh, my gosh, this sounds really dangerous, right, they're using radium bromide, and this is pretty dangerous radioactive material. So, for those of you that don't know radium is extremely radioactive, even in the range of radioactivity, and one of the major problems with it is that if it does get in your body, the radium is treated as calcium in your body. So, you can imagine what happens as it gets deposited into your bones, which is not the ideal situation after a long day in the lab.

So, this is really a pretty dangerous situation that's always interesting to point out. He got this from Marie Curie -- you can imagine they used the postal service, I'm not sure how else they would have transferred it to each other. So, it really brings up some issues. The first thing I did when I heard that is actually look up to see how, in fact, Ernest Rutherford did die in 1937, and you'll be happy to know, it actually wasn't from radiation poisoning or from bone cancer, so that's really good that that worked out okay for him, and that he did get to, sort of safely, at least end his life before the radiation ended it.

But it's really interesting the studies that he did do with radium bromide, and he was studying the alpha particles. And what was known about alpha particles at the time is that they were these charged particles and that they were very heavy. Does anyone know more than what Rutherford knew at the time, what alpha particles actually are? Yeah, good. So, they're actually helium atoms, helium ions. And this wasn't really important for the studies, it didn't matter that didn't know what they are, but it's nice to kind of know now -- that we do know what they were using. And he was doing quite a few studies with them.

One experiment that he was doing is detecting the number of particles that were being emitted by this radium bromide as a rate, so he would measure the number of particles per minute that the radium bromide was emitting. And what he used was a detector here, so he here could detect how many particles were hitting this detector. He had actually developed this detector with a postdoc by the name of Hans Geiger. Does that name ring a bell?

STUDENT: Geiger counter.

PROFESSOR: Um-hmm, a Geiger counter. So, this, in fact, is my very schematic representation of a Geiger counter. For those of you don't know what that is, it's simply an instrument that counts radioactive particles in the air, and now that you're at MIT, you'll all have a chance to see one first hand, if you're ever in any of the labs, especially in the chemistry or bio labs. As carefully as people work with radioactivity here, and using often much, much safer radioactive materials than radium bromide, and using them, and special hoods, and having special procedures, they still do a lot of checking with these Geiger counters to make sure everything's safe. You'll actually see a man walking around with one, sometimes in the halls, just kind of like this -- you hear that click, click, click. That's a good sound, it means low levels of radiation. He'll walk by your hood, so click by your hood -- I always get a little nervous when he walked by my hood, I don't know why, I never worked with radioactive material. But I was convinced I'd hear the click-click-click-click-click, which is what tells you you're in trouble. So, I've never heard the click-click-click-click-click, and we might bring a Geiger counter in here some time later in the semester so we can check all of you out, and hopefully we won't hear any when we do that either.

So, one thing that he discovered with this detector initially, and he was the first to discover this, is that radioactive material, including radium bromide, have a characteristic rate that they emit, radioactive decay. So basically they're decaying at a constant rate, which means, of course, that you can figure out how old things are by seeing how much they've decayed. So, he was really the first person to discover you could do this, which was used to make the first somewhat close approximation of the age of the earth. So that's a pretty exciting set of experiments he did. But one thing that he wanted to do specific to understanding the atom, and using these alpha particles, these heavy-charged particles, was to test if this Plum Pudding model actually fit what he could observe.

So, what he did was he first recorded the count rate of radium bromide before it's going through any kind of a plum pudding atom, and he found that it had a count rate of 132,000 alpha particles per minute were being detected by this Geiger counter. He then set up a situation where he put a very, very thin piece of gold foil right in what would be in the stream of the alpha particles. So, it was only 10 to the negative, 9 meters thick, so about one nanometer, so that's really thin, it's thinner than a strand of hair. So you can imagine, we actually don't need to think of it as a piece of gold foil, it might be easier to think of it as a couple of layers of atoms. So basically he's trying to put some atoms in the way of the alpha particle.

And what he would expect is if this Plum Pudding model is true, nothing's really going to happen to the particles, right, they should go straight through, because if they hit an electron, those are so small. We figured out the mass is so tiny that it shouldn't really deflect them very much. And, of course, all that's left is this positive pudding. So that's not going to do anything either. And what he found when he did this experiment, was that the count rate with still 132,000 counts per minute.

So, what he could conclude thus far was that this was really consistent with the Plum Pudding model. All of his heavily-charged alpha particles were going right through this thin layer of gold atoms.

So, you might think that he would stop his experiments here, and maybe he would have, but as I mentioned, he did have a postdoc working with him by the name of Geiger. He also had an undergraduate, we could say maybe even a UROP working with him, and this was by the name of Marsden was the name of this UROP. And Rutherford realized, you know I have these two people that are very excited to work on this project, I don't need to spend time doing it. Maybe it's not the best way for me to spend time looking to see if I can find any bounced-back particles since all the particles are accounted for. But, you know, this undergraduate's very eager to do it, let's let him have a try. And something you might find in your UROP experience is you have a unique advantage as an undergraduate, which is that there's not a lot of pressure to actually make a huge discovery or necessarily accomplish a great amount. You have a little more pressure in grad school, but sometimes that means when you're an undergrad your advisor will decide to put you on projects that maybe when you look at them seem a little bit silly.

So this project was, let's see if we can detect any alpha particles by making a detector that swings around. So, some people might say, why are we doing this? We know we started with a 132,000 alpha particles. We detected a 132,000 alpha particles. What are we even looking for? We have to build this whole new detector, is this really the best use of my time? As an undergrad, you don't have to worry about it, you're just worried about learning. You can take these big risks of time, and if at the end of the day there's nothing to detect, you still know how to build a detector. So, keep that in mind if you're not over-the-top excited about the prospects of some of your research. You might be surprised at what you find out.

And this is exactly what happened with Marsden who discovered that when he shot the alpha particles at the gold foil, he detected something on his detector that click, click, click went a little bit faster. So, what he detected was that there were 20 alpha particles per minute. Does that sound significant? It depends, right? So hopefully, the first experiment he did, which I know that they certainly did do was maybe it's just background noise, right? So, they took away that gold foil and said is just the alpha particles hitting it some other way? And no, it wasn't. When he took away the gold foil, the count rate went down to zero.

If he switched from gold to let's say iron, he also tried platinum, a number of different foils, he found that they count rate, it still was 20 alpha particles per minute.

So, this is an absolutely outstanding discovery, even though, if we think about it, what is the probability that this happened, how often did this happen? It actually almost happened not at all. We can figure out exactly what the probability of this backscattering was just by dividing the count rate of the number particle that were backscattered divided by the count rate of the incident particles. So essentially, we just have 20, and our 20 is divided by 132,000. So, we end up with a not so large probability of 2 times 10 to the negative 4.

But still, we can't even overstate how exciting this discovery was. Rutherford, the advisor here, he had a lot of good things happen in his life, as I mentioned. He was the person responsible for being able to first date the age of our earth. That's a pretty nice thing. He was also married, he had a child, which I hear is very nice, very exciting, also. But yet, when he saw this one single experiment from this undergraduate, he described this as the most incredible event that had ever happened to him in his life. So, this was a pretty big deal. We won't tell his daughter. And he gave a very good analogy in saying, "It was almost as incredible as if you'd fired a 15 inch shell at a piece of tissue paper, and it came back and hit you."

And that really illustrates what's happening here, because if we think of the Plum Pudding model, it's essentially this very thin film, right, there's nothing that should hit if we send alpha particles through it. But what we actually have is that something's bouncing back. So, what happened is Rutherford needed to come up with a new model for the atom with several interpretations that came out of these experiments, and some of these interpretations were that, of course, we now know that these gold atoms, they must be mostly empty, and the reason that we know that they must be mostly empty is because all but 20 of these 132,000 particles went all the way through. So they weren't hitting anything, we're dealing with mostly empty space.

But he also realized that when they did hit something, what they hit what unbelievably massive, but also that that mass was concentrated into this very, very small space. So eventually, this is what we have come to call the nucleus of an atom. And the nucleus name was used as an analogy to the nucleus of a cell, so in some ways that makes it easier to see the connection, but I think it can also be a little bit confusing for maybe 7th graders that are learning both at the same time, that this nucleus acts very different from a nucleus in a cell, although, of course, there many of them in the nucleus of a cell.

There are some other things that Rutherford was able to figure out. One is the diameter of the nucleus, and that turns out to be 10 to the negative 14 meters. If we think about the size of a typical cell -- excuse me, now I'm getting confused about nuclei. If we think of the size of a typical atom, we would say that would be about 10 to the negative 10 meters. So, we can see the diameter of a nucleus is absolutely smaller, really concentrating that mass into a very small space.

So, you might be asking how did he actually figure that out? We'll do the calculation ourselves. In fact, we'll do the whole experiment ourselves, minus the radioactivity in just a minute, so we'll be able to answer that question for you. He also figured out that the charge of the nucleus was a plus ze. This makes sense intuitively as well, because z is just the atomic number. So, let's say we have an atomic number of 3, that means we have 3 electrons, so we better hope to get our neutral atom that we have a charge of plus 3 in the nucleus.

So, I mentioned at the beginning, while he was working with this radium bromide, that I was very relieved to see that it did not kill him to do these experiments. However, I think I will share with you that the cause of his death was, in fact, related to his research here, even though it was a little more tangled up. So, what happened, of course, after he discovered the nucleus, not surprising, he won a Nobel Prize for this -- I would hope that he would. And in addition to winning a Nobel Prize, he was also knighted, which was a nice bonus for someone born in England, that's a great thing to happen to them. The problem that he ran into is at some point a little bit later in his life, he had a hernia which was a pretty standard case, but what he was going to need was an operation on it. And the glitch came that at least at the time, if you were a knight, you could only be operated on by a doctor that was also titled.

So, Rutherford had a little bit of waiting to do for that doctor to show up, and it turns out the wait was too long, and he actually passed away because he discovered the nucleus and got a Noble Prize and became knighted. So, it's still dangerous. If that opportunity comes up for you, maybe you want to check into the policies of how that works with the doctor situation now. Hopefully they've cleared it up a little bit.

So, what we want to do now is see if we can understand how this backscattering experiment worked. So, we will do our own backscattering experiment. And we'll ask you to imagine a few things. First is that we have this mono layer of gold particles. So let's see if Professor Drennan is able to help us out here. Oh, great. So, that is her daughter, Sam that you see strapped to the chest, and Dr. Patti Christie helping us out here.

All right. So, we'll move this up to the front in just a minute, but I'm going to explain how this experiment works, and we'll do the calculation first before the excitement breaks out. But I'm sure you can easily see how these styrofoam balls could, in fact, be a mono layer of gold nuclei. We have 266, as some of you might know who saw me counting ping-pong balls the other day in office hours. We have 266 ping-pong balls, and we need someone, hopefully you, to be some radioactive material that are going to be emitting these ping-pong balls. And when the time comes, in just a minute, I'll ask the TAs to come down and hand these out very quickly to you, so we can do this experiment.

But first, let's go through how we're going to determine what Rutherford determined, which was he was interested in knowing, which we said what the diameter of the nuclei were. So, we're going to do the same thing and figure out the diameter of these styrofoam balls here, and we can do it by using the relationship of how many backscatter. So, if we think about the probability of backscattering, which is the exact same thing that we saw Rutherford calculate, using the 20 divided by 132,000. But in our case, the probability of backscattering is going to be the number of balls that backscatter, and that's going to be divided by the total number of ping-pong balls. So, do you remember what that was?

STUDENT: 266.

PROFESSOR: 266. Good information retention. All right. So, we have the probability here. So, in terms of the number of balls scattered over the total, we can also relate the probability to the area of all of those nuclei divided by the total area that the atoms take up.

Right, this makes a lot of sense, because if the entire atom was made up of nuclei, then we would have 100% probability of hitting one of these nuclei and having things bounce back. So, here we have the area of the nuclei we'll figure out adding those all together versus the space of all of the atoms put together.

So, not only did Professor Sayer, who's in the Chemistry Department who put together this contraption for all of you, not only did she magnify the size of these gold nuclei, but she actually had to smoosh all of these atoms closer together then they normally would be. If, in fact, a gold nucleus was this size here, we would need to use another lecture hall in order to find a place to put this nucleus right here. This is a little bit of a tricky experiment, so we decided we'll just smoosh it all in, and we'll actually be able to account for it, because we'll take into account the area of all of those atoms.

I think this board does not like to go by itself. All right. So we can figure out what that is, the area of all of the nuclei is going to be the number of nuclei times the area per nucleus, and we're going to talk about the cross-section here to keep it simple. And all of that is divided by the area of the atoms, which is 1 . 39 meters squared, measuring that space there.

So, the number of nuclei, if we were to sit and count these as well, is 119. So, we'll multiply that by just pi, r squared, to get that cross-section, and divide all of that by 1 . 39 meters squared.

So, what we have here is a relationship that can tell us what the probability of backscattering is, but what we want to pull out, since we can experimentally measure what the probability is, what we need to pull out is the radius or the diameter of these nuclei, so we can just, instead of solving for p, we can switch it around and solve for the radius. So, that's going to be equal to the probability raised to the 1/2 times 6 . 098 times 10 to the negative 2 meters.

So, actually, just for discussion sake, it makes a little more sense for us to talk about the diameter, so that's just twice the radius. So, once we figure out what our probability of backscattering is, we'll just raise that to the 1/2, and we'll multiply that by 12 . 20 centimeters.

All right. So now all we have to do is figure out this probability of backscattering. We know we need to divide by 266, but what we need you to help us with is to figure out this top number here and see how many particles are going to backscatter. So, if the TAs can come up and quickly hand out 1 particle to everyone. And a few people will need to throw 2, if you feel like you have particularly good aim.

PROFESSOR: So, as you're getting your ping-pong balls -- do not throw them yet. Let me explain to you what constitutes a backscatter event. So, it'll be considered a backscatter event if your ping-pong ball hits one of the nuclei. It's not going to be a backscatter event if your ping-pong ball hits the frame or these strings, or the top part.

So, in a few minutes, not now, we're going to ask you to stand up, and you can kind of come over more toward the center of the room if you want, and aim your ping-pong ball at the lattice here, follow the ping-pong ball with your eye, and discover, watching it, whether it's a backscatter -- it hits one of the nuclei and bounces back towards you, or if it goes through, and also if your ping-pong ball doesn't land anywhere in the vicinity of this at all, then keep that in mind. And then at the end of the experiment we'll ask you what happened to your ping-pong ball, and you'll let us know, and we can calculate the number of backscatter events.

Are there any questions before we get started? Raise your hand if you don't if you don't have a ping-pong ball yet. Any questions before we get started? All right. So, we'll come around and get ping-pong balls to the rest of you. Those of you who have your ping-pong balls can now begin the experiment.

[EXPERIMENTING]

PROFESSOR: All right. Any last shots? There we go. All right. So, it looks like we were a little bit successful, I saw some backbouncing. We were going to have a clicker slide on how many bounced back, but it looks like we're having a little technical difficulty with that. So, what I'll ask is can you stand up if you had your particle bounce back?

All right, so let's count how many we have here. So, 13 backscattered. TAs, if you can maybe pick up these ping-pong balls for me. I'm sure it would be very amusing if I fell, but I'd rather not.

All right, so, we have 13 divided by 266. All right, MIT students, who has a calculator on them? Actually, I should probably do it as well, so I know I'm hearing correctly. So, are you getting . 0489 or so? All right. So, we've got our probability. We can go ahead and plug that in, take the square root of it, multiply it by 12 . 2. What are you getting for your diameters? Yup, that's what I got, too. All right. So, we have 2 . 70 for our diameter, and that's in centimeters. So, we actually did a pretty good job here. It turns out that the diameter is actually 2 . 5 centimeters. So, good job, experiment well done, plus we were not exposed to radioactivity, which is a bonus.

So, this is exactly how Rutherford did discover that these particles were present and made this new model for the atom that we now know has both a nucleus, and we know the size, and also has electrons. So, to finish up today, we won't get through all of it. But the next thing we can actually talk about is now that we know we have an atom that has a nucleus, let's say somewhere in the center, and it has electrons around it, thinking on our most simple example, which is hydrogen, we have a nucleus and an electron that have to hang together in the atom in some way, and we need to think about well how can we describe how atoms behave, and specifically, how do we describe how any single atom stays together where the two are associated, but at the same time they don't immediately collapse into themselves. So, what we can do is try using the classical description of the atom and see where this takes us.

So, if we think about the force that occurs between a positively and a negatively charged particle, what we have is essentially a Coulomb force, so we can describe this as a force of attraction. We can use the Coulomb force law to explain this where we can describe the force as a function of r. So, let's think about what we're saying here. We're describing the force that's holding these two particles together, and it's related to the charge of each of the particles, where e is the absolute value of an electron's charge. So, an electron has a charge of negative e, we've written here, and the nucleus has a charge of positive e. And then we have r, which is simply the distance between the two charges. And what we see is that the force is inversely related to the distance between the two charges. And we can simplify this expression as saying negative e squared over 4 pi, epsilon nought r squared. Epsilon nought is a constant, it's something you might see in physics as well.

Essentially for our purposes here, you can just think of it as a conversion factor. What we need to do is get rid of the Coulomb tag that we have -- that's how we measure our electron charges -- charge, and so we use this epsilon nought quite often, this permativity constant of a vacuum to make that conversion. And I'll just point out here also, this is a conversion factor you'll use quite frequently -- many of you, quite on accident, will memorize it as you use it over and over again. But I do want to point out that you don't have to memorize it for any exams in this class, we will give you a sheet that has all the needed constants that you're going to use on there, so save up that brain space for other information. ah

So, we can use Coulomb's force law, and we can think about these different scenarios. So, when you come in on Monday, we're going to start off, you can think for the weekend -- you probably only need to think for a second about what happens when r goes to infinity, but that's where we'll start on Monday. And let me just suggest to all of you also, that you get those problem sets started this weekend. You should absolutely finish, at least through part a this weekend, and save part b for next week. So, have a great weekend.