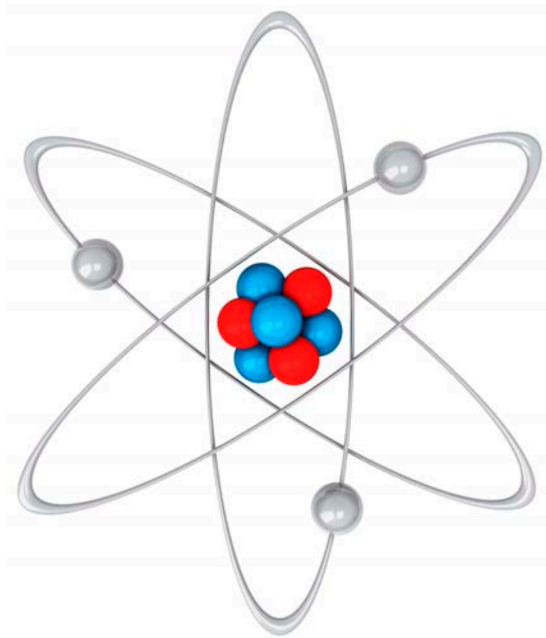
Chapter 2: Our Model of the Atom

Introduction

Have you seen a drawing like the one on the right? You probably have. It is an out-of-date **model** of the atom. What do I mean by "model"? In science, we often have to study things that are either too small to see or too complicated to describe easily. In these situations, we use a model to help represent what is being studied.



This is a model of an atom

Model – A representation of a process or an object in nature that allows us to more easily understand it.

So a model is part of a theory. A theory attempts to explain why something in the natural world behaves the way it does, and a model makes the theory easier to understand. The model on the right is one way to picture what an atom looks like. It makes the overall theory of the atom easier to understand, because it illustrates the important aspects of the atom. Now as I said, it is an out-of-date model, but it is still a commonly-used illustration. What does it mean and why is it out of date? That's what you will learn in this chapter.

Learning That Atoms Are Made of Different Particles

If you have been studying the history of science, you have probably learned that scientists have been talking about atoms since before Christ was born. Over time, it became accepted among scientists that all **matter** is made up of atoms. Matter includes everything you see around you, except for light, which is pure energy.

Matter – A general term for any physical substance

At the beginning of the 19th century, John Dalton came up with a working hypothesis of atoms that seemed to explain all the observations that had been made regarding how matter behaves. He used it to predict something that had never been observed before, and the prediction was confirmed, so his hypothesis became a theory which is now called **Dalton's Atomic Theory**.

In Dalton's atomic theory, atoms were thought to be **indivisible**, which means they could not be broken down into smaller parts. However, as time went on, some experiments indicated that wasn't true. Before I explain those experiments, do one of your own:

Experiment 2.1: Making Charges

Supplies:

- A balloon
- An aluminum can (like the kind soda comes in)
- Clean, dry hair or a wool sweater or blanket

This experiment works best when the air is not very humid.

Instructions:
1. Blow up the balloon and tie it off.
2. Lay the aluminum can down on its side on a flat table, countertop, or floor. The countertop should not be made of metal, and it should be level enough that the can doesn't start rolling on its own.
3. Hold the balloon about 30 centimeters (12 inches) to the right of the can, slightly above the table, countertop, or floor.
4. Slowly move the balloon towards the can and watch the can carefully. Continue to do this until the balloon touches the can. Did the can move before being touched by the balloon?
5. Put the can back where it was and rub the balloon vigorously in your hair. If your hair isn't clean and dry, use someone else's hair or a wool sweater or wool blanket.
6. Repeat steps 3 and 4, making sure the part of the balloon that you rubbed in your hair is facing the can. What is different now?
7. Once again, rub the balloon in your hair.
8. Once again, repeat steps 3 and 4, but when the can starts to roll towards the balloon, pull the balloon back so that the can continues to roll. See how long you can make the can "chase" the balloon.
9. Clean up your mess.

What did you see in your experiment? The can should not have moved the first time you brought the balloon close to it. It might have moved away from the balloon once the balloon touched it, but that's just because it was pushed by the balloon. What happened after you rubbed the balloon in your hair? The can should have started rolling towards the balloon. In the next part of your experiment, you should have been able to "pull" the can with the balloon, as long as you kept the balloon close to but ahead of the can.

How can we explain these results? Well, scientists have known for a long time that electricity comes in two forms, which Benjamin Franklin labeled as **positive** and **negative** charges. It was also known that opposite charges attract one another, while like charges repel one another. When the can and balloon were first brought close to one another, neither of them had an overall electrical charge, so they were neither attracted to nor repelled from one another.

Once you rubbed the balloon in your hair, you caused it to develop an electrical charge. Specifically, you caused it to develop a negative charge. When the negatively-charged balloon was brought near the can, the can was attracted to it. This means that there must have been positive charges in the can, and those positive charges were attracted to the balloon. But where did those positive charges come from? You charged the balloon, not the can.

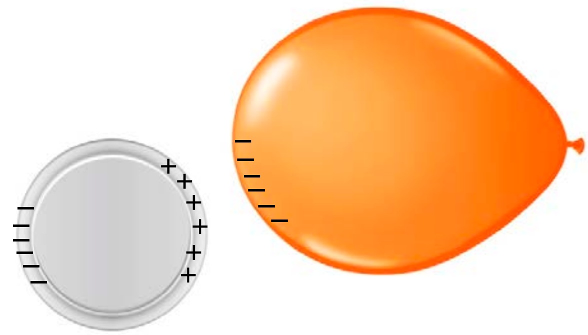
Believe it or not, the positive charges were always there in the can, but there were an equal number of negative charges that cancelled them out. The same was true for the balloon before you rubbed it in your hair. Since there were equal numbers of negative and positive charges in both, there was no overall charge on either, so the can and balloon weren't attracted to one another. When you rubbed the balloon in your hair, it picked up some extra negative charges from your hair, so the balloon became negatively charged overall.

What happened when the negatively-charged balloon was brought close to the can? The negative charges in the balloon repelled the negative charges in the can. The can is made of aluminum, which is a **conductor**.

<u>Conductor</u> – A substance through which electrical charges can move

Since the negative charges were repelled by the balloon, and since they could move in the can, they moved *away from* the balloon. Now remember, there were equal numbers of negative charges and positive charges in the can. When the negative charges moved away from the balloon, the parts of the can from which they moved now had more positive charges than negative charges, and the area toward which they moved had more negative charges than positive charges.

The result of all this charge-moving was that the side of the can that was farther from the balloon developed a negative charge, while the side of the can that was closer to the balloon developed a positive charge. In the end, the situation looked something like the drawing on the right. The positive charges on right side of the can were attracted to the negative charges in the nearby balloon. What about the negative charges in the can that were repelled by the negatively-charged balloon? They were farther away from the balloon, so their repulsion wasn't as strong as the attraction of the positive charges. That means there was an overall attraction to the balloon, so the can started rolling towards the balloon.



The negative charges on the balloon (right), forced the negative charges in the can (left) to move so that the positive charges were close to the balloon and the negative charges were far from the balloon.

So what does this have to do with atoms? Remember that the can, balloon, and your hair are all made up of atoms. When you rubbed the balloon in your hair, the balloon picked up negative charges. Where did those negative charges come from? *They came from the atoms in your hair*. In the same way, the negative and positive charges in the can were in the atoms of the can. Since the negative charges in your hair went to the balloon and the negative charges in the can were able to move away from the balloon, what does that tell you about the negative and positive charges in an atom? It tells you that they can be separated!

Of course, that means Dalton's Atomic Theory is wrong, at least when it comes to atoms being indivisible. As your experiment indicates, atoms are made up of positive and negative charges, and under the right circumstances, those positive and negative charges can be separated from one another. Scientists figured this out near the end of the 19^{th} century, but in a much more dramatic way, which you will learn about the next time you do science.

Comprehension Check

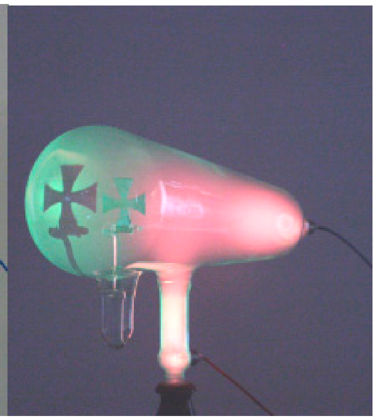
2.1 After you rubbed the balloon in your hair, did your hair have a charge? If so, was it positive or negative?

2.2 Suppose the can had been made of something that charges cannot move in. If you did everything exactly the same as you did in your experiment, would the can have been attracted to the negatively-charged balloon?

Measuring Ratios

In the late 1800s, a scientist named **William Crookes** (krooks) had been experimenting with glass containers that are now called "Crookes tubes." He removed most of the air in the glass container and replaced it with a small amount of a specific gas. He also put a metal object in the tube,

Crookes tube without the electricity turned on.



Crookes tube with the electricity turned on.

like the cross shown in the pictures on the left. He then hooked one part of the container to the positive side of a battery and another part of the container to the negative side of the battery. If the battery was strong enough, it made the gas and the glass container glow, as shown in the picture on the near left! The pinkish color you see comes from the gas glowing, while the greenish color you see comes from the glass glowing. That's pretty cool, isn't it?

Since the metal object cast a "shadow" on the glass, Crookes realized that something must be flowing through the tube, and the metal object was blocking that flow. He thought that the electricity from the battery had changed the gas to a previously-undiscovered state of matter, which became known as "cathode rays." However, British scientist **Joseph John (J.J.) Thomson** had a different idea. Based on the fact that a magnet bent the path of these mysterious "cathode rays," he thought that they might be electrically charged. So he ended up putting metal plates inside the tube, charging one plate positively and the other one negatively. There were several problems he had to overcome to get the system working properly, but eventually, he was able to show that the particles always bent in the direction of the positive plate. What did that tell him? Remember, opposite charges attract, so Thomson concluded that the particles must be negatively charged.

This is a photo of J.J. Thomson

But how much negative charge do these particles have, and what is their **mass**? If you don't recognize the term "mass," it is a measure of how much matter is packed inside an object.

Mass – A measure of how much matter exists in an object

As you might expect, the more matter there is in an object, the heavier it is. Thus, mass and weight are related to one another. The larger an object's mass, the heavier it is. But please remember that even though mass and weight are related to one another, *they are not the same*.

Mass and weight are related to one another, but they are not the same.

You will learn *a lot more* about the difference between mass and weight when you take chemistry in high school. For right now, just realize that the more mass an object has, the heavier it is.

Most of the time, scientists measure mass by putting an object on a scale. However, these negatively-charged particles existed only in a Crookes tube and only when the electricity was on. There was no way Thomson could use a scale to measure their mass. Nevertheless, Thomson was able to measure something. Using the charged plates, he was able to measure the charge-to-mass **ratio**.

Ratio – The relationship determined by dividing the first quantity by the second quantity

In other words, he couldn't measure the charge that these cathode rays had. He also couldn't measure their mass. However, he could measure the number that is produced when the charge is divided by the mass. How could he measure the ratio when he couldn't measure the two numbers? Perform the following experiment to see that you don't necessarily need to measure two quantities in order to measure their *ratio*.

Experiment 2.2: Measuring Mass-To-Volume Ratio

Supplies:
- Honey
- Water
- Vegetable oil
- A tall, clear glass
- A coin (like a penny)
- An ice cube
- A wooden match
- A grape (optional)

Instructions:
1. Pour enough honey in the glass so that it forms a layer about 2.5 cm (one inch) thick.
2. Slowly add water to the glass, allowing it to slide down the side of the glass. Add enough to make a layer that is about the same thickness as the honey.
3. Repeat step 2 with vegetable oil.
4. Set the glass down and look at its contents. You should see three distinct layers: a layer of honey at the bottom, a layer of water in the middle, and a layer of vegetable oil on the top.
5. Carefully drop the coin in the glass.
6. Carefully drop the ice cube in the glass.
7. Carefully drop the match in the glass.
8. If you have a grape, drop it in the glass as well.
9. Look at the contents of the glass again. Where are the objects that you dropped into the glass?
10. Clean up your mess. Use lots of dish soap and hot water to get the honey and vegetable oil out of the glass. Also, squirt some dish soap down the drain so that the honey and oil go down the drain easily.

What did you see in the experiment? As you were told after you added the three liquids to the glass, you should have seen three distinct layers: a honey layer at the bottom, a water layer in the middle, and an oil layer on top. These layers formed for two reasons. First, the three liquids don't mix well. As a result, they all tend to stay separate from one another. Second, the honey was at the bottom not because you added it first, but because it had the highest **density**.

<u>Density</u> – The mass-to-volume ratio of a substance

Since volume is a measure of how much space an object occupies, density is the mass of a substance divided by how much space it occupies. Think about what that means. When matter is tightly packed in a substance, there is a lot of mass in a small volume. As a result, when you divide mass by volume, you get a large number. The larger the density, the more tightly-packed the matter is in a substance.

You might have already learned that things will float in water if they weigh less than an equal volume of water. A ship is made of metal, but it can still float, because a ship-sized sample of water weighs a lot more than the ship weighs. Since mass and weight are related, we can also say that something will float in water if its mass is less than the mass of an equal volume of water. Do you see where I am going with this? Density tells us the ratio of the mass to the volume, so something will float in water if it is less dense than water! In fact, we can say something even more general than that:

If a substance has a lower density than a liquid, it will float in that liquid.

Well, the density of water is lower than that of honey. As a result, the water floated on the honey! That's why the honey stayed at the bottom of the glass, and the water stayed above the honey. Now think about the vegetable oil. What can you say about its density? Its density is lower than that of water, because the vegetable oil floated on the water!

This coconut floats in water because it is less dense than water.

Think about that for a moment. Did you measure the mass of the oil, water, or honey? No. Did you measure the volume of the oil, water, or honey? No. However, by seeing how they ended up floating on one another, you now know that the density of vegetable oil is lower than the density of water, and the density of water is lower than that of honey. Well, remember what density is. It is the mass-to-volume ratio of a substance. So without measuring the mass or the volume, you were able to determine which had the greatest mass-to-volume ratio (honey), which had the lowest mass-to-volume ratio (vegetable oil), and which had a mass-to-volume ratio that was in between (water).

You can use the same reasoning when it comes to the solid things you dropped into the glass. The coin sank all the way to the bottom, so it has a density (mass-to-volume ratio) greater than all the liquids, including honey. If you had a grape, it should have sunk through the oil and water, but it should have floated on the honey. That means it has a density that is greater than that of water, but less than that of honey. Similarly, the ice cube had a density that is greater than that of vegetable oil but less than that of water. The wooden match had a density that is less than everything else in the glass.

Comprehension Check

2.3 Object **A** has twice the mass of object **B**, and object **B** has twice the mass of object **C**. Which object weighs the least?

2.4 Object **A** has a higher density than water, while object **B** has a lower density than water. Which object floats in water, and which sinks?

The Beginning of a Model for the Atom

In the previous experiment, you measured the mass-to-volume ratio of different things without measuring either the mass or the volume. J.J. Thomson did something similar with cathode rays. Rather than measuring their mass-to-volume ratio, he measured their *charge-to-mass ratio*. However, just like what happened in your experiment, he didn't have to measure the charge and the mass and then divide them. He would have loved to do that, but he couldn't. However, his experiment at least allowed him to measure the charge-to-mass ratio, and he found out that it is very large.

So what did that tell Thomson? Think about it. You get the charge-to-mass ratio by dividing the electrical charge by the mass. Since the charge-to-mass ratio was large, that means the mass is smaller than the charge. Why? Well, when you divide 2 by 100, what do you get? You get a small number, 0.02. That's because you are dividing a number by something that is larger. However, suppose you divide 100 by 2. What do you get then? You get a big number, 50. So when you divide a number by something smaller than itself, you get a big number. Since the charge-to-mass ratio that Thomson got was large, that told him the mass is smaller than the charge.

Remember that the charged plates had already told Thomson that the charge on these cathode rays was negative, so Thomson concluded that atoms must be composed of negatively-charged particles and positively-charged particles. The electricity had separated those charged particles from each other, producing the particles whose charge-to-mass ratio he calculated. He originally called those negative particles "corpuscles," but they were later named **electrons** (ih lek' trahnz). As a result, we say that J.J. Thomson discovered electrons, the negative particles in an atom.

Before I continue, I want to share with you something that Thomson said in a speech that he gave to the British Science Association:

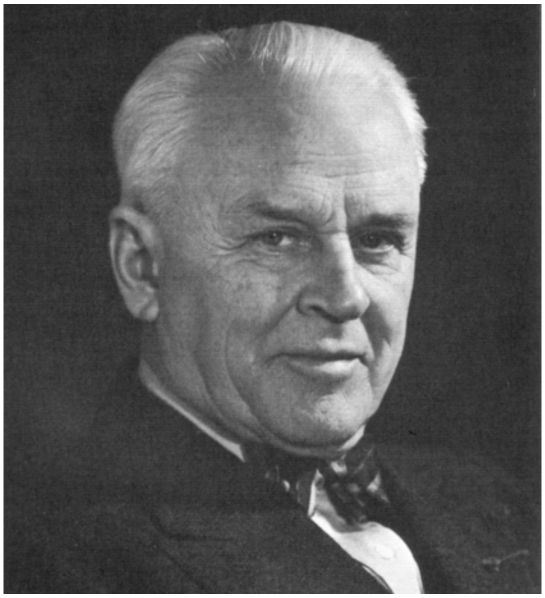
> As we conquer peak after peak we see in front of us regions full of interest and beauty, but we do not see our goal, we do not see the horizon; in the distance tower still higher peaks, which will yield to those who ascend them still wider prospects, and deepen the feeling, the truth of which is emphasized by every advance in science, that "Great are the Works of the Lord." (*University of California Chronicle*, Vol 19, p. 34, 1917)

Thomson was a devoted Christian, spending time in prayer and Bible study every day. He understood that as scientists, we are studying the handiwork of God.

With the discovery of the electron, scientists had to come up with a model for the atom that included two parts: a positive part and a negative part. After all, Dalton's Atomic Theory said the atom had no parts; it is indivisible. Thomson showed that idea was wrong. Now remember, a model is a representation of something that helps us understand it better. So what could represent an atom that had positive and negative charges which could be separated? Sir William Thomson, known as Lord Kelvin, claims that he was puzzling over this one day while eating his "pudding." He was an Englishman, so what he calls "pudding," we would call "dessert." One such dessert is **plum pudding**, which is pictured on the right. It is a cake-like dessert with raisins embedded in it. Lord Kelvin suggested that atoms might look like plum pudding. The atom has a basic structure of positive material that would be like the cake-like substance of plum pudding, and the raisins would represent electrons, embedded in the positive material.

This is one version of plum pudding, which inspired Kelvin's model of the atom.

J.J. Thomson (who was no relation, even though both have "Thomson" in their names) was in complete agreement with Kelvin on this model of the atom, and he began to promote it as well. Lord Kelvin was near the end of his life at this point and had already made many contributions to our understanding of nature, especially when it comes to the study of energy. Because of his reputation, the **plum pudding model** was considered the best model of the atom in the early 1900s. Lord Kelvin was not only a great scientist, but he also encouraged scientists to see God in what they were studying. In 1903, he said this in a short address at the University College in London: "Do not be afraid of being free thinkers. If you think strongly enough you will be forced by science to the belief in God, which is the foundation of all Religion. You will find science not antagonistic, but helpful to Religion." (*The Nineteenth Century and After: A Monthly Review*, Vol. 53, James Knowles, ed., Sampson, Lowe, Marston, and Company 1903, p. 1069)



This picture of Robert Millikan was taken about five years before he wrote "Science and Religion."

Despite J.J. Thomson's experiments, there were some who were still skeptical of the idea that atoms had individual parts which can be separated. However, a few years later, an American scientist named **Robert A. Millikan** produced more evidence for the electron's existence, and along the way, he figured out the electrical charge that it possessed. Millikan was the son of a minister and was fully convinced that God had designed everything that scientists study. When he was in his fifties, he wrote a paper for the *Bulletin of the California Institute of Technology* (March 1922 edition) entitled "Science and Religion." In it, he wrote, "I have never known a thinking man who did not believe in God."

Millikan designed an elegant experiment, which is a bit too complicated to explain in detail. However, it involved taking a tiny drop of oil and giving it a negative charge. He then used a device that he made to measure the charge that he had just given the drop of oil. He did this over and over again. While all the drops he used were small (he had to use a microscope to see them), they were of different sizes. Also, the charges that he measured were different. Some oil drops had a lot of charge on them, while others had only a little charge on them.

His experiment was very difficult. In fact, once he had it completely set up and working, it took him 60 days to get reliable measurements for just 58 oil drops! Think about that. He averaged a bit more than a day to get a good measurement for each oil drop. That's how difficult his experiment was. However, he really wanted to learn as much as he could about how things become electrically charged, because he hoped it would help him better understand electrons.

His perseverance paid off. He found something amazing in his results. To understand why it is amazing, think about Experiment 2.1. In that experiment, you rubbed a balloon against your hair (or something made of wool) to give it a negative charge. Think about what would have happened if you had just rubbed the balloon in your hair a couple of times. It might have gotten a little charge, but not much, right? Instead, you were told to rub the balloon in your hair vigorously. That's because you needed a lot of charge on the balloon, and the more you rubbed the balloon in your hair, the more negative charge it would get. So the amount of charge on the balloon depended on how much you rubbed it in your hair: The more you rubbed it, the more charge it would get.

Based on this information, you might think that you could give the balloon any amount of charge you wanted. After all, the more you rub the balloon in your hair, the more charge you can give it. You might expect the same thing for Millikan's oil drops. He could give them any amount of charge he wanted, depending on how long he charged them. However, when Millikan looked at the 58 drops he measured, he didn't see that. Instead, he saw that the charge of each oil drop was a **whole-number multiple** of a specific number: -0.000000000477 electrostatic units. Now don't worry about the number itself or what "electrostatic units" means. Just understand that it represents a specific amount of negative charge. To make things simpler, I will call that number "*e*."

What do I mean when I say "whole-number multiple"? I mean that every charge on all 58 drops was some whole number (like 1, 2, 3, 4, etc.) times the number that I am calling "*e*." In other words, these drops couldn't have just any charge. They could have a charge of 1 times *e* or 2 times *e*, but nothing in between. None of them could have a charge of 1.3 times *e* or 1.8 times *e*. Only 1 times *e* or 2 times *e*. They could also have a charge of 3 times *e*, but they couldn't have a charge of 2.5 times *e*. No matter what charge he measured, it was always a whole-number multiple of *e*.

What could that possibly mean? Well, imagine that I am holding a bunch of coins in my hand and you ask me how much money I have. Suppose I say that I am holding 94.5 cents. Would you believe me? You shouldn't! You should know that in the United States, the smallest coin is a penny, which is valued at one cent. As a result, if I am holding a bunch of coins, their value must be a whole-number multiple of a penny's value. I could be holding 94 cents or 95 cents, but I could be holding 94.5 cents, because one cent is the smallest value for a coin.

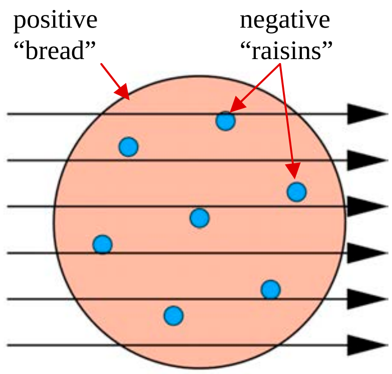
When it comes to a set of coins, then, the value must be a whole-number multiple of one cent, because that's the smallest unit of money that exists in today's U.S. coins. Since Millikan's oil drops had charges that were whole-number multiples of *e*, that means *e* is the smallest unit of charge in creation. Since electrons are the negative parts of an atom, Millikan decided that *e* must be the charge on the electron. When he charged his oil drops, he gave each of them a certain number of electrons, and as a result, they had a charge equal to the number of electrons he gave them times *e*.

Now remember, J.J. Thomson had already measured the charge-to-mass ratio of the electron. With Millikan's measurement of the electron's charge, the electron's mass could be calculated as well! Thus, Millikan's experiment, when combined with Thomson's experiment, taught us two very important things about the electron!

An Improved Model for the Atom

Millikan's experiment provided powerful evidence for the existence of the electron, which meant that scientists had to figure out exactly how the electron fits into the atom. Now, of course, scientists could have simply believed the plum pudding model of the atom, since Lord Kelvin was such a well-respected scientist, and since the man who showed that atoms must have positive and negative charges agreed with him. However, scientists must continually test their ideas to make sure they are reasonable. This is where **Ernest Rutherford** comes in. He had worked for J.J. Thomson, so he had become interested in the plum pudding model. He was also familiar with the work of **Marie Curie**, about whom you might have learned before. She had discovered that uranium spontaneously produced positively-charged particles in a process that she called **radioactivity**. Rutherford studied radioactivity intensely, and he figured out a way to use it to test the plum pudding model of the atom.

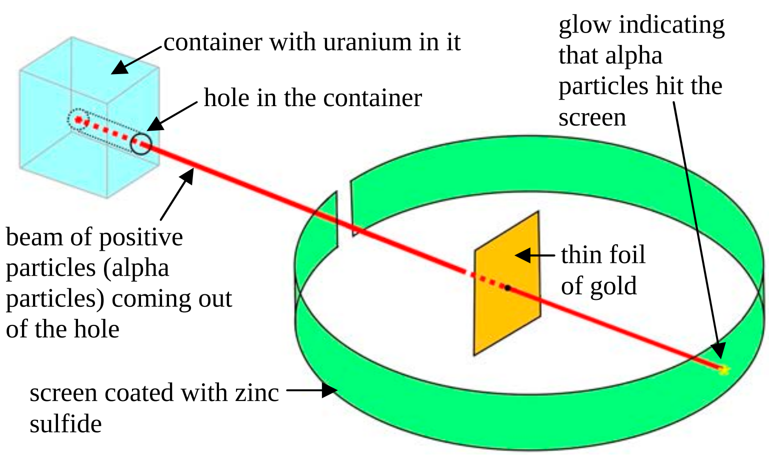
Rutherford had determined that the positive particles produced by uranium, which were called **alpha** (al' fuh) **particles**, could be stopped pretty easily, so he put some uranium in a container that had a single hole in it. The container stopped any alpha particles that didn't travel through the hole, so the container essentially made a "beam" of positive particles. It would be like putting a black cover with a hole in it over a light bulb. The light from the light bulb would be absorbed by the black cover, so the only light that could escape would be the light coming out of the hole. As a result, you would have a beam of light that would travel in the direction that the hole was pointed.



A plum-pudding atom should repel alpha particles (the black arrows) as much as it attracts them, so the alpha particles should travel pretty much straight through the atom.

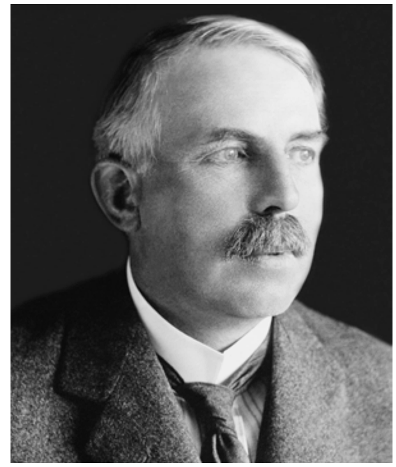
How would this help Rutherford test the plum pudding model of the atom? Well, think about how positive particles should behave in the presence of other charges. They should be repelled by other positive charges and attracted by negative charges, right? So what would happen if the beam of particles were aimed at an atom? If the plum pudding model were correct, the beam of positive particles coming out of the container would be repelled by the positive "bread" but attracted by the negative "raisins" (the blue dots) that are scattered throughout the bread. What would be the overall effect? The beam of positive particles would be repelled and attracted, but if the "bread" and "raisins" were pretty much evenly distributed, the positive particles would be repelled just as much as they would be attracted. The result would be that the beam of particles should travel pretty much straight through the atom, as shown in the illustration on the left. They might bend towards an electron part of the time, but then they would eventually bend the other way because of the other charges. As a result, they might move in bendy paths while in the atom, but by the time they left the atom, those bends should cancel out, and it would look like they traveled pretty much straight through the atom.

Now, of course, actually doing an experiment to test this was pretty difficult. After all, you can't see the atoms, and you can't see the positive particles in the beam. Nevertheless, Rutherford was pretty ingenious. About 40 years earlier, a French chemist had shown that a chemical called zinc sulfide would glow if it were given some extra energy. It turns out that when alpha particles collide with zinc sulfide, they give it the energy that it needs to glow. The glow doesn't last long, but it happens right where the alpha particles hit the zinc sulfide. As a result, zinc sulfide can be used to detect alpha particles.



This is a simplified sketch of the experiment Rutherford designed.

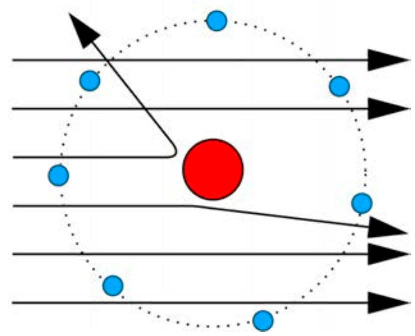
So Rutherford had two of his assistants, Hans Geiger (gy' gur) and Ernest Marsden, set up the experiment drawn on the left. The beam of positive particles would leave the hole in the container, pass through the gap in the back of the screen, and then hit the thin foil of gold. They would then pass through the atoms in the gold foil, and if the plum pudding model of the atom were correct, they should travel in pretty much a straight line. As a result, there would be a single glow on the screen where all the alpha particles would be hitting it.



This is a portrait of Ernest Rutherford.

As is the case with many scientific experiments, the results were unexpected. Yes, there was a lot of glowing exactly where Rutherford, Geiger, and Marsden expected it to be, but to their surprise, they saw glows on other parts of the screen as well. Some of the glows were even *behind* the gold foil, near the gap in the back of the screen! This indicated that some of the positive particles were hitting the gold atoms and bouncing around. Some even bounced almost directly backwards! Rutherford described the results this way: “It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell [a bullet fired from a large, powerful cannon] at a piece of tissue paper and it came back and hit you” (Edward Neville da Costa Andrade, *Rutherford and the Nature of the Atom*, Peter Smith Pub Inc,1964, p. 111).

There was simply no way to explain the results in terms of the plum pudding model. However, Rutherford could explain the results another way. Remember, most of the alpha particles passed straight through the foil, as expected. However, some of the alpha particles bounced around. Suppose the positive and negative charges weren’t evenly spread out around the atom. Instead, suppose the positive charge was concentrated at the center of the atom, and the negative charges were on the edge of the atom. What would happen then?



Placing a concentrated positive charge at the center of the atom explained the results of the experiment. Most of the alpha particles would travel straight, but some would bounce away.

Well, when the alpha particles passed in between the positive and negative charge, they would be attracted by the negative charge and repelled by the positive charges. On average, then, they would behave the same as if they had traveled through a plum-pudding-type atom. However, think about what would happen when the alpha particles passed very close to the positive charge. The positive alpha particles would be repelled by the positive charge strongly, because the closer two charges are, the greater the electrical force between them. The alpha particles would still be attracted by the negative charges of the electrons, but that attraction would be weaker. As a result, the repulsion would “win,” and the alpha particles would be bent away from the positive charge. The closer an alpha particle got to the central, positive charge, the more strongly it would be bent, so that some of the alpha particles would bounce backward, as shown in the illustration on the right.

Now remember, the positive charge at the center attracts the electrons, so the electrons can’t be just sitting there. They would be drawn into the center. Instead, Rutherford realized that the electrons would have to orbit around the center of the atom, just like the planets orbit around the sun. As a result, this model of the atom became known as the **planetary model**. The positive charge at the center of the atoms was called the **nucleus** (new’ klee us).

<u>Nucleus</u> – The center of the atom, where there is a concentrated, positive charge

This model was consistent with the results of the experiment that Rutherford designed, but there were still a lot of questions regarding the specifics of the model. Did the electrons all follow one orbit, as shown in the drawing above? What determined the size of the orbit? Was the nucleus just one big lump of positive charge, or was it made up of a bunch of little charges? Those were questions that needed to be answered!

Comprehension Check

2.5 Objects **A** and **B** have the same charge. If object **A** has more mass than object **B**, which has the higher charge-to-mass ratio?

2.6 You give an object a negative charge by adding 15 electrons to it. If you represent the charge of an electron with "*e*", what is the charge on the object?

2.7 If you were able to completely remove an electron from an atom, would the atom have a charge? If so, would it be positive or negative?

More about the Planetary Model of the Atom

Rutherford was able to do some more experiments to help him understand the planetary model of the atom a bit better. He noticed that different kinds of particles hitting the zinc sulfide screen caused different kinds of glows. For example, if he got hydrogen atoms moving very quickly and slammed them into a zinc sulfide screen, they would cause a different kind of glow than the alpha particles did. He also noticed that if he just let alpha particles travel through the air and into a zinc sulfide screen, he would get glows from the alpha particles as well as glows that looked like hydrogen atoms. He eventually found that if he had the alpha particles travel through pure nitrogen gas, he would get even more glows that looked like hydrogen.

One way to understand this result is that sometimes, an alpha particle would hit the nucleus of a nitrogen atom head-on. When that happened, it would "break" the nucleus, causing part of the nucleus to be kicked out. That kicked-out part is what caused the glow that looked like hydrogen. Well, that would mean that the nucleus of a nitrogen atom is made up of a bunch of hydrogen nuclei (the plural of nucleus).

He proposed, therefore, that the nucleus of every atom was composed of a bunch of particles that were the same as the one particle found in the nucleus of a hydrogen atom. He didn't come up with that reasoning entirely on his own. He had read the work of William Prout, who had been doing experiments nearly 100 years previously. Prout had done some experiments that indicated the mass of every atom is a whole-number multiple of the mass of a hydrogen atom. Once again, that means the mass of any other atom can be gotten by multiplying the mass of a hydrogen atom by a whole number: 1, 2, 3, 4, etc.

Since Rutherford's experiments seemed to indicate that the nucleus of a nitrogen atom had several nuclei of hydrogen atoms inside, it made sense that the nucleus of every atom was composed of the same particles that made up the nucleus of a hydrogen atom. An atom that is heavier than hydrogen, then, has several hydrogen nuclei in its nucleus. Scientists didn't want to call the things inside other atoms "hydrogen nuclei," so Rutherford suggested a couple of names to use instead. Eventually, the word **proton** (pro' tahn) was accepted.

So by this time in history (the early years of the 1900s), the atom was thought to consist of two particles: electrons and protons. The protons are packed together in the nucleus of the atom, and the electrons orbit around the nucleus. Protons are positively-charged, and electrons are negatively-charged. One proton's positive charge perfectly cancels out one electron's negative charge. Atoms start out with equal numbers of protons and electrons, so atoms don't have any overall electrical charge. However, if you remove electrons or add electrons, an electrical charge can be generated.

While all of this might sound perfectly reasonable, some students have a hard time understanding why electrons would orbit around a positively-charged nucleus. After all, as you have already learned, opposite charges attract one another. Why don't electrons just travel *towards* the nucleus? The nucleus is attracting the electrons, and the electrons are attracting the nucleus. Why don't they just pull together? You could ask the same thing about the planets and the sun. The sun attracts the planets, and the planets attract the sun, because gravity is an attractive force between masses. Why don't the planets just crash into the sun? Perform the following experiment to find out.

Experiment 2.3: Centripetal Force

Supplies:
- A ball (It can be as small as a golf ball or as large as a baseball.)
- A plastic bag that can hold the ball but still have extra room
- A 50-cm (20-inch) length of string that is strong enough to hold several times the weight of the ball that you have chosen

Instructions:
1. Put the ball in the bag.
2. Tie one end of the string securely around the opening of the bag so the string keeps it tightly closed, as shown on the right. That way, the ball cannot fall out of the bag.
3. Lay the ball on the floor as you see in the picture.
4. Hold the free end of the string with one hand, and put your other hand on the floor right next to where you are holding the free end of the string.
5. Pull on the string with the hand that is holding it. What happens to the ball? It moves in the direction you are pulling it, getting closer to the hand that you put on the floor. Not all that exciting, is it?
6. Once you are appropriately dressed for the weather, go outside with the string, bag, and ball.
7. Find a place that is far from any windows or other breakable objects.
8. Wrap the free end of the string around your index finger twice, and then make a fist.
9. Press your thumb against the string that is wound around your index finger. This should hold the string tightly in your hand.
10. Raise your hand above your head and start twirling the ball so that it travels in circles around your head. Concentrate on how the string feels as the ball twirls. Do you feel it getting tighter as the ball travels faster?
11. Once again, making sure you aren't anywhere around something that is breakable, open the hand that is twirling the ball, spreading your fingers apart from one another. This will release the free end of the string.
12. Notice what happens.
13. Pick up the ball, bag, and string. Untie the string and put everything away.

Most likely, you weren't surprised by anything that happened in the experiment. However, the results explain why the planets orbit the sun instead of crashing into it. They also explain why electrons orbit around a positively-charged nucleus without crashing into it.

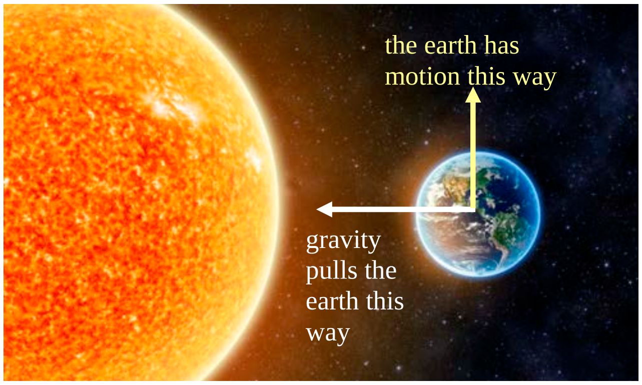
Think about the first part of the experiment. You pulled on the string, and the ball moved in the direction you pulled it, causing the ball to get closer to the hand that you had put on the floor. Why? When you exert a force on something, you change its motion. If someone throws a ball to you and you catch it, you exert a force on the ball. The force you exert changes the motion of the ball, causing it to stop. In the first part of your experiment, the ball was at rest. When you exerted a force on it by pulling on the string, that force changed the motion of the ball, causing it to move toward the hand that you had on the floor.

But what happened in the second part of the experiment? You got the ball twirling in a circle. The string got tight, because it was pulling on the ball. In other words, the string was exerting a force on the ball. That force was directed toward your hand, because that's where the end of the string was. However, the ball didn't move toward your hand. So even though the string was exerting a force on the ball, and even though that force was directed toward your hand, the ball did not move any closer to your hand.

What was the force doing? Think of what happened in the last part of the experiment. When you let go of the string, it could no longer exert a force on the ball. What happened? The ball flew away. As long as the string could exert a force on the ball, it kept that from happening. However, as soon as the string could no longer exert a force, the ball flew away. So the force that the string was exerting on the ball was changing the motion of the ball, *keeping it from flying away*. In essence, the string was pulling the ball toward you, but the only thing that pull was able to accomplish was to keep the ball from flying away. It couldn't bring the ball any closer to you. We call that a **centripetal** (sen trip' uh tul) force.

Centripetal force – A force that acts on an object moving in a circle, directed toward the center of the circle

Now don't get this confused with something called "centrifugal force." That's completely different. In fact, it's not even a real force, but that's something you'll probably learn about in another course.



The gravity of the sun acts as a centripetal force, keeping the earth in circular motion. It is not strong enough to pull the earth into it. It is only strong enough to keep the earth from moving farther away. (Note: the earth is not really that close to the sun.)

The gravitational force that keeps the planets orbiting the sun in the solar system is like the force that the string exerted on the ball. It is a centripetal force. It pulls the planets toward the sun, but because the planets are all moving, it can't actually pull them in. All it can do is change their motion so that they travel around the sun. The attraction between the nucleus and the electron is similar. The electron is moving, and its attraction to the nucleus changes the motion but is not strong enough to pull it in. It can only change the motion enough to keep the electron from moving farther from the nucleus.

Comprehension Check

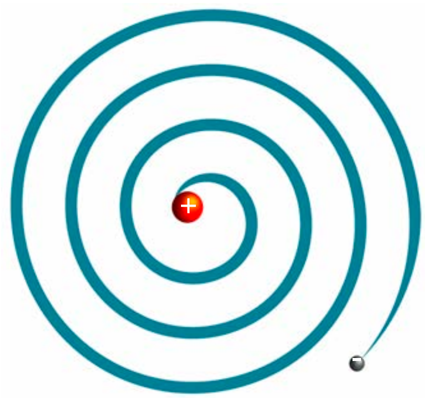
2.8 Suppose an atom consists of three protons and three electrons. Draw what it would look like according to Rutherford at this point in time.

Atoms and Light

While it is reasonable to compare an electron orbiting a nucleus to a planet orbiting the sun, there is one *big* difference between the two situations. Planets don't have electrical charge, but electrons do. This actually presented a really big problem for the planetary model of the atom. About 50 years prior to Rutherford's experiment, James Clerk Maxwell had shown that whenever a charged particle's motion changes, it must emit some light. Often, that light isn't visible, but that's just because human eyes aren't designed to see it. It is still there, and it is emitted by the charged particle whose motion is changing.

Well, if a charged particle emits light, what happens to the particle? It *loses energy*. Think about it. We know light is a form of energy, and we also know that energy can't be created or destroyed. So, when a charged particle emits light, the energy that makes the light must come from the charged particle. Thus, the charged particle loses that energy. What happens to something that loses energy while it is in motion? It *slows down*.

Think about what this would mean for the electron orbiting the nucleus. As the force that pulls the electron toward the nucleus changes the motion of the electron, the electron will emit light. That means the electron will lose energy. Therefore, it will slow down. But what's keeping the attraction between the proton and electron from pulling the electron towards the nucleus? It's the *motion of the electron*. Each time the electron slows down, then, the attraction between the nucleus can pull the electron a bit closer to the nucleus. As time goes on and the electron slows down even more, the attraction will pull it even closer to the nucleus. In the end, then, the electron should not be able to orbit the nucleus in a circle. It should spiral into the nucleus, as shown in the illustration on the right.

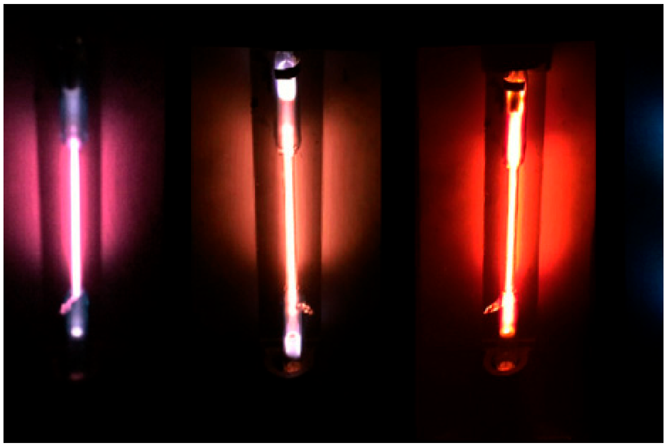


According to everything that was known at the time, the planetary model should not be stable. Instead, the electron should spiral into the nucleus.

Now please understand that Maxwell's conclusion about charged particles emitting light when they change their motion was well-understood and well-documented. It was considered a law of nature. However, the model of the atom that was consistent with Rutherford's experiment required an electron to continually experience changes in its motion but not emit any light at all. So what did Rutherford do? He did something that is pretty common in science. He just assumed that there was something we didn't quite understand about electrons and that they must behave differently in the atom than the charged particles that work according to Maxwell's law.

Was there any reason for him to assume that? Sort of. It was the only way he could make sense of his gold-foil experiment's results. The fact that he knew Maxwell was right about charged particles emitting light did not shake the confidence he had in his explanation of his gold-foil experiment, so he just decided to continue to work on his model of the atom, putting aside the fact that it seemed impossible based on the science that was known at the time.

That might sound really odd to you, but it happens all the time in science. Often, while scientists are working on their models, they have to "put aside" the science that contradicts those models. This doesn't mean they are ignoring the science that is known at the time. It is just an admission that science can't explain everything, so we must sometimes pursue options that seem to contradict accepted science. Of course, the option being pursued might be wrong because it contradicts the accepted science of the day, or it might be right because the accepted science of the day is wrong. In this particular case, the accepted science of the day did not apply to atoms (you will see why later), so Rutherford was right to "put it aside" while working on his model.



Different elements emit different colors of light when they are excited by electricity.

Even though atoms don't continually emit light as would be expected from the science that was known at this time in history, scientists had found that they could force atoms to emit light. For example, if you fill a tube with a gas and connect it to a source of electricity, the gas will glow. Interestingly enough, however, the glow is different for every gas. Consider the picture on the left. Three tubes are filled with three different gases, each of which is an element (from left to right: mercury, hydrogen, and neon). They are each connected to a source of electricity, and they each glow with a different color.

Scientists didn't have a problem understanding the basics of what's going on in the picture. The electricity is giving the atoms in the gas energy. Those atoms absorb the energy for a while, becoming more energetic. However, after a while, they release that energy in the form of light. So in a situation like the picture above, electrical energy is being converted by the atoms into radiant energy (light). The fact that atoms excited by electricity emit light wasn't a problem for scientists of this time to understand. However, the *details* of that emitted light were puzzling.

This is a picture of Niels Bohr and his future wife a year before he got his Ph.D.

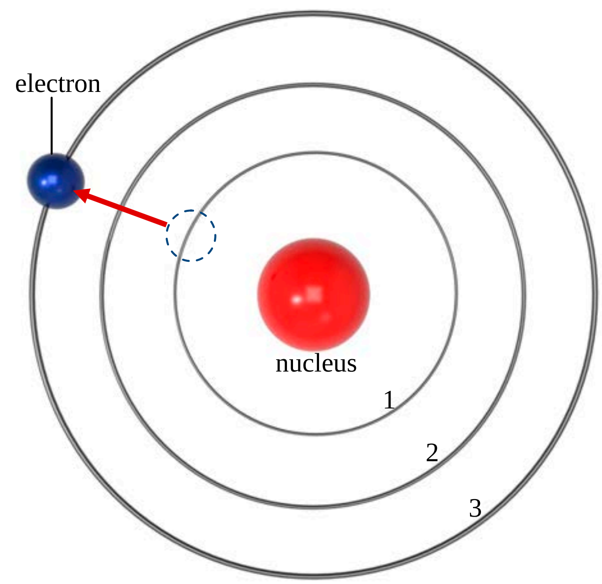
Why, for example, did each element emit its own color of light? Also, when you look at the light in detail, you find that only specific energies of light are emitted, and each element has its own specific energies of light. That made no sense. Why should atoms emit light in this way? Why can't they emit light with a wide range of energies?

This is where **Niels Bohr** (neelz boor) came in. He started working with Thomson shortly after earning his Ph.D. in physics, and while working there, he met Rutherford. Bohr became intrigued with Rutherford's planetary model of the atom and decided to see if he could modify it in some way to explain things like the light emitted by excited atoms. He ended up finding a way to do that, but he had to make a pretty crazy assumption in order to get the job done. You might be surprised to hear about a scientist making a crazy assumption, but that actually happens quite often. Sometimes, it takes some pretty crazy assumptions to make sense of experiments.

So what was this crazy assumption? He decided there were certain "special" orbits in which the electron can orbit the nucleus without emitting any light. The electron can be in any one of those orbits, but *it cannot be anywhere in between*. Why is that crazy? Well, Bohr was working with the "planetary model" of the atom, so think about the planets. Each planet is in a specific orbit around the sun, but it doesn't have to be there. It could be closer to the sun or farther from the sun, as long as it was moving at a different speed.

Remember, the gravitational attraction between the sun and a planet changes its motion, but not enough to pull it in – just enough to keep it from flying off. If a planet moved a bit slower, it could be in an orbit a bit farther from the sun, because the gravitational attraction needed to keep it from flying off wouldn't need to be as large. In the same way, if it moved faster, it could be closer to the sun, because the gravitational attraction needed to keep it from flying away would need to be greater. For every possible speed of the planet, there would be a corresponding orbit.

That's not the way Bohr's atom worked. The electrons could only be in specific orbits, so the electrons could only have specific speeds. Think about it. Do you know any other moving thing that is restricted to specific speeds? If I throw a ball, the harder I throw it, the faster it goes, right? A professional baseball player is supposed to be able to throw a baseball with speeds of greater than 90 miles per hour (40 meters per second). However, if he throws it with a bit less energy, he can make it go a bit slower. If he throws it with a bit more energy, he can make it go a bit faster. Suppose it was not possible for him to do that. Supposed he could only throw a baseball at speeds of 70 miles per hour, 80 miles per hour, 90 miles per hour, or 100 miles per hour, but nothing in between. That's crazy, right? Nevertheless, that's what Bohr had to assume.



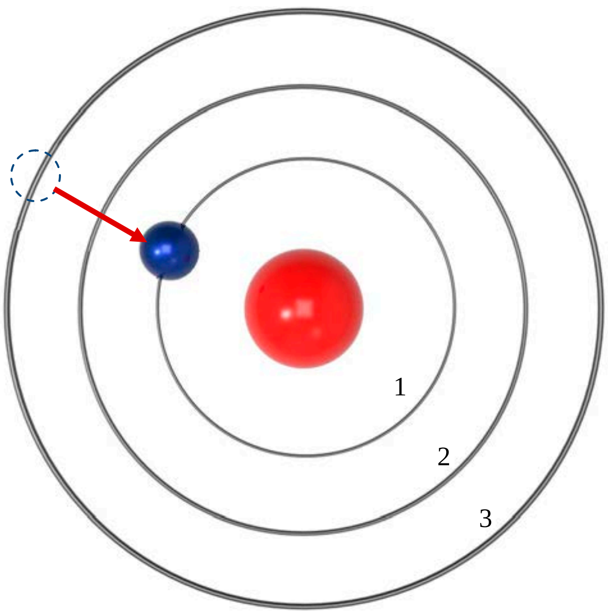
In the Bohr model, an atom absorbs energy by having an electron "jump" from an orbit close to the nucleus (like orbit 1) to an orbit far from the nucleus (like orbit 3).

Bohr decided that when the electron is in an orbit close to the nucleus, it doesn't have a lot of energy. When it is far from the nucleus, it has more energy. If you give an atom energy (by passing electricity through it, for example), then, the electron can jump from an orbit close to the nucleus to an orbit farther away. But since the electron can only be in specific orbits, and since each orbit has its own energy, that means the electron would have to absorb a specific amount of energy. The drawing on the right illustrates this model, called the **Bohr model**, and how an atom absorbs energy.

Remember, with this model, there are only certain orbits that the electron can be in. In the drawing above, they are shown as orbits 1, 2, and 3. The electron can be in any of those orbits, but it cannot be anywhere in between. So let's assume it starts in orbit 1. When electricity is used to excite the atom, the electron can absorb energy, *but it can't absorb just any amount of energy*. It can only absorb the specific amount of energy needed to get from orbit 1 to another allowed orbit. In the drawing, it is moving from orbit 1 to orbit 3, so it has to absorb the exact amount of energy needed to get from 1 to 3. It can't absorb a bit more or a bit less. It has to absorb that specific amount of energy.

This means that atoms can't just absorb just any amount of energy. They have to absorb specific amounts of energy – the specific amounts needed for the electrons to go from one specific orbit to another specific orbit. In other words, energy has to be absorbed in "packets." A "packet" of energy is called a **quantum** (kwan' tuhm) of energy. Because Bohr's model makes the assumption that energy must be absorbed in "packets," we say that it is a part of **quantum theory**.

Now remember, Bohr wanted to explain the light that is emitted by excited atoms. Can you see how his theory does this? Think about the electron on the previous page. It is in a high-energy orbit. How can it get back to a lower energy orbit? It can emit some light. Light is energy, and if the electron emits light, it loses energy. But think about the energy of that light. What would it have to be? Suppose the third orbit (the one the electron is in) has a certain energy, E_3. Suppose the electron wants to get back to the first orbit, which has an energy of E_1. What energy of light would have to be emitted to do that? The difference between those two energies, $E_3 – E_1$. If the electron emits that amount of energy, it goes from the third orbit to the first orbit.



In the Bohr model, an electron in a high-energy orbit (like orbit 3) can go back to a low-energy orbit (like orbit 1) by emitting light that has energy equal to the difference in the two orbits' energy.

There is another way that the electron can emit light and get down to the first energy level. Think about what that way would be. It could first jump from the third orbit to the second orbit, and then it could go from that second orbit to the first orbit. In other words, it could take two small steps down to the first orbit instead of one big step. What would it have to do? It would have to emit *two different energies of light*. First, it would emit light of energy $E_3 – E_2$, the difference in energy between the third and second orbit. Once in the second orbit, it would then emit light once again, but this time, the energy would be $E_2 – E_1$.

So the electron has "choices," as far as how to get rid of that energy, but no matter what choices are used, in the end, only specific energies of light can be emitted, and those energies are equal to the difference in energy between the orbit it starts in and the orbit it jumps down to. When Bohr applied his model to the hydrogen atom, he made some assumptions about what orbits would be possible for the electrons, used some commonly-accepted equations from physics, and was able to calculate the specific energies of light that an excited hydrogen atom could emit, using the choices the electron had. Guess what? Those energies were exactly the ones that had already been observed for hydrogen!

To scientists of the day, this was powerful evidence that there was a lot of truth to the Bohr model of the atom. After all, no one had any explanation as to why excited atoms emitted only specific amounts of energy. Nevertheless, with the Bohr model, the light emitted by excited hydrogen atoms could not only be explained, but its energy could be calculated! Now don't let Bohr's success make you forget that his assumptions are really crazy. Bohr himself admitted it. In fact, he once wrote, "Anyone who is not shocked by quantum theory has not understood it" (*Journal and Proceedings of the Royal Society of New South Wales*, Vol 116-119, p. 51). As time went, it became somewhat of a running joke. Late in his life, he heard another scientist describe a new theory and said, “We are all agreed that your theory is crazy. The question that divides us is whether it is crazy enough to have a chance of being correct.” (F. J. Dyson, "Innovation in Physics," *Scientific American,* Vol 199, pp. 74-82, 1958)

Comprehension Check

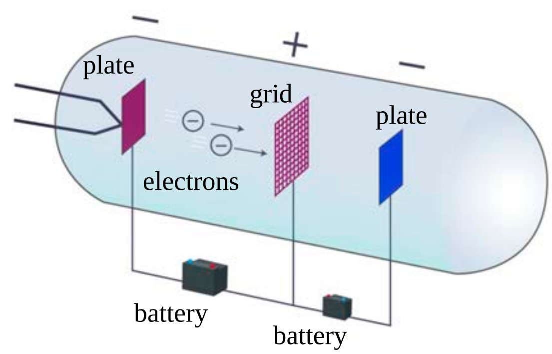
2.9 Suppose a charged particle that is in motion absorbs light instead of emitting it. Would the charged particle lose energy, gain energy, or end up with the same energy?

2.10 An electron in an atom jumps from an orbit that is close to the nucleus to one that is farther away. Did the atom absorb energy or lose energy?

2.11 An atom is emitting light. Is its electron moving closer to the nucleus or farther from the nucleus?

Testing the Crazy Assumption

The idea that an atom’s electrons can only absorb specific amounts of energy was so crazy that scientists had to test the idea. **James Franck** (frahnk) and **Gustav Hertz** (goo stahv’ hurtz) came up with an ingenious way to do just that. Their experiment, now known as the **Franck-Hertz experiment**, is drawn on the right. They used a tube with a small amount of mercury gas in it. The tube had a plate connected to the negative end of an electrical source, and then, near the other end of the tube, there was a grid connected to the positive end. Electrons would travel from the plate to the grid, being attracted by the positive charge on the grid. This attraction would make them speed up, moving faster and faster as they approached the grid.



The Franck-Hertz experiment consisted of a tube of mercury gas, two plates, and a grid.

By the time they reached the grid, they would be moving quickly, so they couldn’t just stop when they reached the positive charge. Since the grid had lots of holes in it, many electrons would pass through the holes and keep going. A short distance from the other side of the grid, there was a plate that was connected to a small negative charge. The electrons would be attracted to the positive charge on the grid and repelled by the small negative charge on the plate, so they would start to slow down. Electrons without much energy would stop and move back to the positive grid, but electrons with a lot of energy would slam into the negative plate despite being repelled by it, because the repulsion just wasn’t strong enough to stop them. Franck and Hertz measured how many electrons slammed into the grid by measuring the electrical current that was produced.

Now think what should happen as Franck and Hertz increased the voltage on the grid in their experiments. The higher the voltage, the larger the positive charge on the grid. The larger the positive charge on the grid, the faster the electrons should be moving when they reached the grid. Thus, the higher the energy they should have. The higher the energy, the more electrons should slam into the plate at the end, so the higher the current. As the voltage increased, then, there should be an increase in electrons that hit the plate at the end.

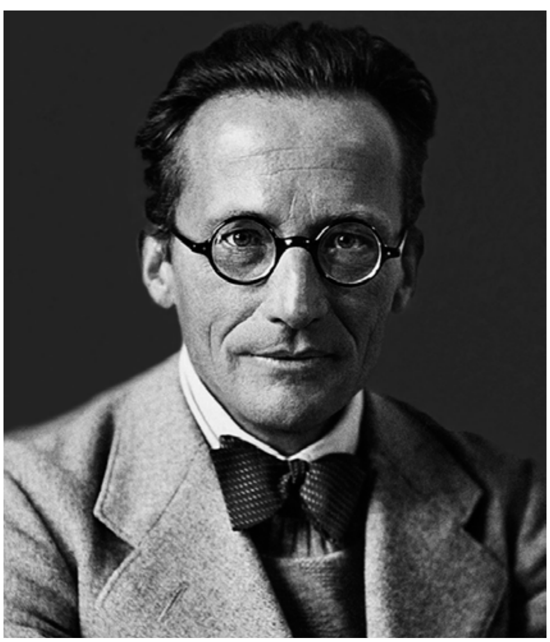
At first, that's exactly what happened. As the voltage increased, the electrons that hit the plate on the end increased. However, once the voltage passed 4.9 volts, the number of electrons hitting the plate decreased significantly. After decreasing, it would increase for a while, but at 9.8 volts, it decreased again. Every multiple of 4.9 volts saw another decrease in the number of electrons on the plate.

What could possibly explain these results? Well, a decrease in electrons on the plate would indicate that electrons were losing energy as they moved. Otherwise, they would make it to the plate. What could cause them to lose energy? Remember that there is mercury gas in the tube. That means there are mercury atoms in the tube. The electrons could collide with those atoms. If they lost energy during those collisions, that would reduce the number of electrons that hit the plate at the end. But the reduction occurred only at specific voltages: multiples of 4.9 volts. What does that tell you? A specific voltage would produce a specific energy in the electrons. Thus, they could only lose energy in collisions with a mercury atom if they had a specific amount of energy.

In the end, electrons were colliding into mercury atoms, but when those electrons didn't have the right amount of energy, the mercury atoms couldn't take any energy away from them. However, when the electrons had the right amount of energy, the mercury atoms could take it away from them, reducing the number of electrons on the plate. In other words, the Franck-Hertz experiment showed that mercury atoms could only absorb specific amounts of energy, which confirmed Bohr's crazy assumption!

So even though Bohr's assumption made no sense, it not only led to an atomic model that could explain why excited atoms emit light at only specific energies, but it also was confirmed for mercury atoms. As a result, the Bohr model became the leading model of the atom at this time in history (1914). Of course, there was still a lot more to learn.

Explaining One Crazy Assumption with Another



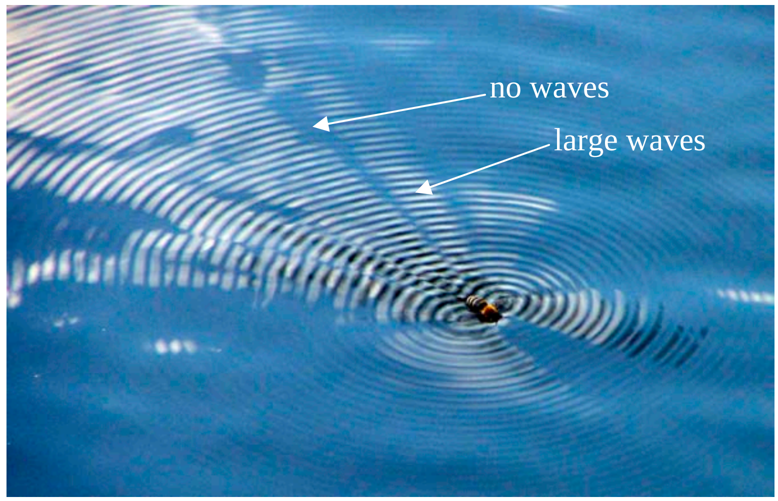
This is a photograph of Erwin Schrödinger.

Since the Franck-Hertz experiment demonstrated that atoms can only absorb specific amounts of energy, scientists really wanted to understand why. Yes, it comes from the fact that electrons can only occupy certain orbits around the nucleus, but why is that? If it has the right speed, a planet can be in any orbit around the sun. Why can an electron only be in specific orbits around the nucleus?

Erwin Schrödinger (shrow' ding er) decided to think about the problem in a completely different way. Thus far, scientists had been thinking about electrons as tiny particles. However, they behaved a bit like light. For example, go back and look at the Crookes tube images on page 18. Notice how the cross-shaped object casts a "shadow" on the glass. You could say the electrons traveling through the glass act like rays of light, and the cross stops those rays of light, making a shadow. Schrödinger decided to treat electrons in an atom as if they were like light.

How would that help? Well, light can be thought of as waves (like waves on water), and waves behave differently from particles. For example, they can interfere with other waves. Look at the picture below. The bee is moving its wings trying to get out of the water. One wing is making one set of waves, and the other wing is making another set of waves. When the two sets of waves meet, they start overlapping with one another, making a strange pattern. In some places, the waves are really large. In other places, the waves are small or even non-existent! The resulting pattern is called an **interference pattern**, and it exists because waves can add to and subtract from each other.

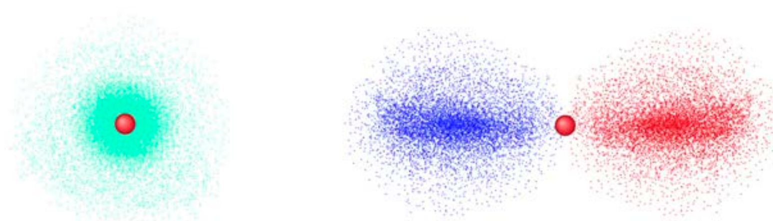
Perhaps the reason electrons can be in only specific orbits is because they act as waves, setting up an interference pattern like the one on the right. Wherever there are high waves, there are electrons. Wherever there are no waves or tiny waves, there are no electrons. Now once again, this is a crazy assumption. Why would electrons behave like waves? They have mass and charge, like a particle. Nevertheless, as you have seen, sometimes a crazy assumption can produce some pretty amazing results, so Schrödinger decided to make the crazy assumption and see where it went.



The two sets of waves made by the bee's wings produce an interference pattern, where there are larger waves and no waves.

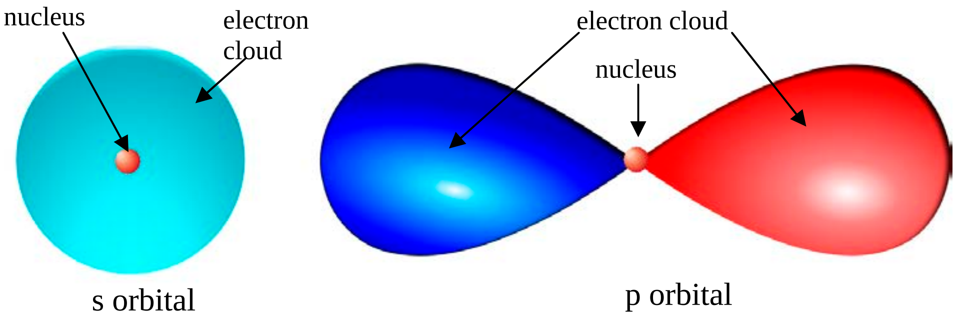
It turns out that the behavior of a wave can be described by a particular equation, which is unsurprisingly called a "wave equation." Schrödinger applied a wave equation to an electron in an atom, and he found that electrons would, indeed, produce interference patterns in an atom. However, those interference patterns didn't result in simple circles, like the orbits in the Bohr atom. Instead, they resulted in fuzzy "clouds," such as the two shown below.

In each of the drawings, the sphere at the center is the nucleus. Each dot represents a place where the electron might be at a certain time. So if you could watch an electron move in an atom, you would see it flitter about in a seemingly random way, but if you marked each spot in which you saw the electron with a dot, you would eventually see a pattern. In some cases, the electron would move about within a sphere, like you see on the left. In other cases, it would move about in a more complicated shape, like the one you see on the right. According to Schrödinger's wave treatment, then, electrons really don't orbit the nucleus. Instead, they flit around the nucleus, constrained to be in fuzzy shapes that are sometimes called **electron clouds**.



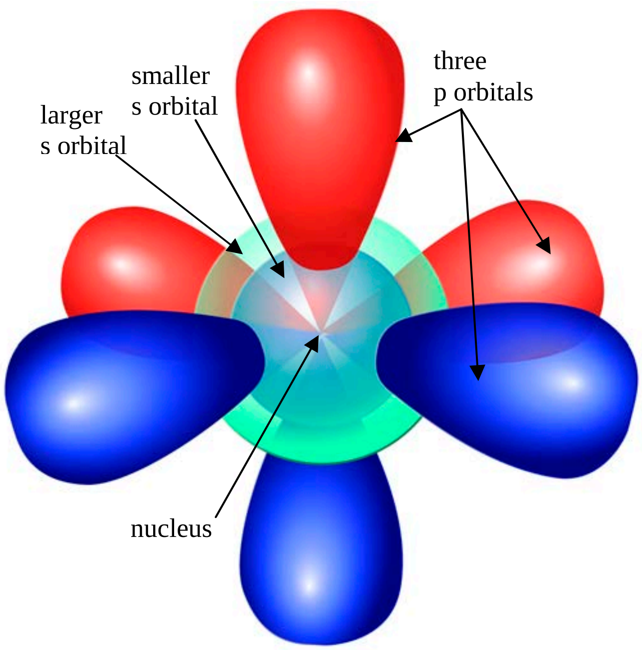
In Schrödinger's view, electrons don't orbit the nucleus. Instead, they move about in clouds that have geometric shapes.

While drawings like the ones on the previous page are helpful, they don't really show off the fact that the clouds in which the electrons travel are three-dimensional. Also, the atom is pretty complicated, so it is best to simplify the clouds a bit to make them easier to understand. So scientists usually represent electron clouds with well-defined shapes like the ones shown on the left. We call these shapes **orbitals**, and they are simplified versions of the clouds drawn on the previous page. A spherical cloud is called an **s orbital**, while two-lobed clouds are called **p orbitals**.



Electron clouds are called orbitals, and they are represented with well-defined shapes, even though they are more like the fuzzy clouds on the previous page.

That's not the end of the story. There can be many s orbitals and p orbitals in an atom, each with different sizes depending on the energy of the electrons inside. In addition, the more interestingly-shaped orbitals, like the p orbitals, can be oriented differently around the nucleus. As a result, atoms can look really, really complicated. For example, the drawing below is the way Schrödinger would view a neon atom. Notice that it has a total of five orbitals. There is a smaller spherical orbital (a smaller s orbital) inside a larger spherical orbital (a larger s orbital). In addition, there are three two-lobed orbitals (three p orbitals), each of which is oriented in a different direction.



This is the modern view of how electrons travel around the nucleus in a neon atom.

This is essentially the modern view of how electrons move around in an atom. We call this the **quantum-mechanical model** of the atom. We don't know that it is correct, because we can't see atoms to find out if they look like this. However, we can say that lots of experiments have been done to test this model, and the experiments work out just as the model predicts. Thus, it is hard to believe that the model isn't at least somewhat correct.

We aren't quite done, however, because at this point in time (1926) scientists didn't understand something very basic about the nucleus of the atom. You will see what that is in the next part of the chapter.

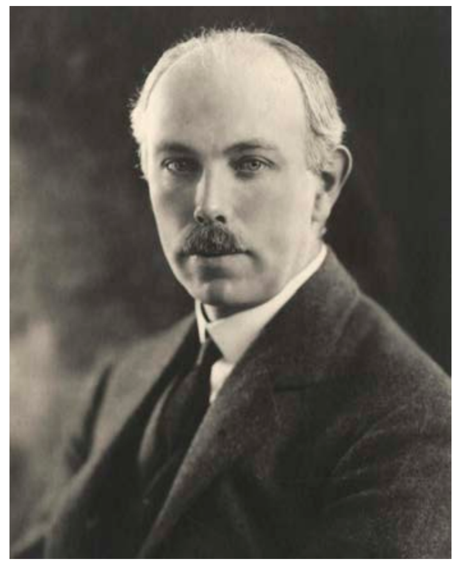
Comprehension Check

2.12 In some college laboratories, students do the Franck-Hertz experiment with neon in the tube instead of mercury. When the experiment is performed that way, would you expect the decreases in current to happen at multiples of 4.9 volts or some other number of volts?

2.13 Look at the picture of the neon atom on the previous page. Suppose you measure the energy of an electron in the smaller s orbital, and then you measure the energy of the electron in the larger s orbital. How would the energies compare? (**HINT**: Think about electron energy in the Bohr model.)

There's Something Else Going on Here

I need to take you back in time a bit now so that you can see how scientists came to understand the nucleus of the atom a bit more clearly. Remember how J.J. Thomson discovered the electron. He used a Crookes tube and measured the charge-to-mass ratio of the negative particles that were traveling through the tube. From that, he decided that electricity was separating atoms into negative particles (which we now call electrons) and positive particles (which we now call protons). People could see the effect of the negative particles traveling through the Crookes tube. The negative particles made the gas and the glass light up. But what about the positive particles? If the electricity was separating the atoms into negative and positive particles, there had to be positive particles streaming through the Crookes tube as well.



This picture of Francis William Aston was taken shortly after World War I.

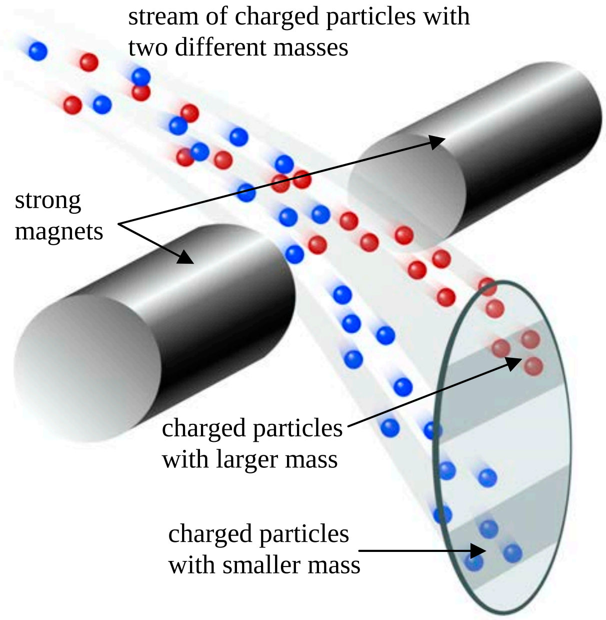
Thomson wanted to learn about the positive particles as well, but they were harder to do experiments with. Thomson was able to master the experiment to measure the charge-to-mass ratio of the negative particles, but doing the same with the positive particles seemed to give him lots of contradictory results. As a result, he invited **Francis William Aston** to work with him. Aston had become well-known for making modifications to already-existing scientific instruments that improved their performance significantly. Thomson thought that perhaps with Aston's help, he could finally figure out a way to measure the charge-to-mass ratio of those positive particles.

Aston started working with Thomson in 1909, but even with his help, the results of the experiment still didn't make any sense. When working with neon, for example, it seemed that they kept getting two different charge-to-mass ratios. The negative particles had only one charge-to-mass ratio, so they expected the same for the positive particles. That's not what their experiment was telling them. Aston started working on more improvements to the experiment, but then World War I broke out. Aston started working for the Royal Aircraft Establishment (he and Thomson were in England), studying how to make airplanes more resistant to bad weather conditions.

Aston didn't return to his study of positive particles until 1919, but once he did return, he made rapid progress. He modified the experiment he and Thomson were doing in an ingenious way, making the first **mass spectrograph** (spek' truh graf).

Mass spectrograph – An instrument that measures the mass of individual atoms or molecules

Today, mass spectrographs are used by many scientists around the world.



This is a very simplified drawing of a mass spectrograph. Since a magnetic field bends the path of charged particles based on their mass, you can use it to determine the mass of atoms or molecules.

How do you measure the mass of a single atom or molecule? You can't put a single atom on a scale! However, you can see how the particles are affected by magnets. If a charged particle is moving in a magnetic field, it will travel along a curve, and the bend in that curve depends on its mass. If you set things up properly, you can send a stream of charged particles into a magnetic field, and the particles will follow different curves, depending on the mass. If you detect them after they have traveled through the magnetic field, you will see separate beams of particles, each with a different mass, as shown in the drawing on the left.

Now please understand that mass spectrographs are a lot more complicated than this, and Aston's was especially complicated. Remember, he and Thomson thought their initial experiments gave them strange results, so he wanted to make sure that his mass spectrograph was very precise. When he put neon in his mass spectrograph, he saw two distinct signals, indicating two distinct masses. This was just like the results in the previous experiments, and they were perplexing. After all, neon was thought to be an element. According to Dalton's Atomic Theory, the atoms of an element are supposed to be identical in every way. Nevertheless, his mass spectrograph told him that there were actually two forms of neon: a heavier form and a lighter form.

He then did the same experiment with chlorine, which was also thought to be an element. He saw two strong mass signals from chlorine, and one weaker signal that might have indicated a third mass. So there were up to three masses for the element chlorine. Based on his experiments, he concluded that Dalton was wrong about atoms of a given element being identical. In the end, it seemed that elements can be made up of atoms that have different masses. He confirmed this by studying many other elements and once again, finding that each of them were composed of atoms with more than one mass. He described them using a term that had already been suggested, **isotopes** (eye' suh tohps).

<u>Isotopes</u> – Atoms in the same element that have different masses

So even though Dalton's Atomic Theory was very important, it was shown to be wrong on two key points. First, as you learned at the beginning of this chapter, atoms are not indivisible. Second, as Aston demonstrated, atoms of a given element are not identical. They can have different masses.

But what explains these isotopes? How could an element be composed of atoms that have different masses? The answer came as a result of some experiments with radioactivity. Remember the alpha particles that Rutherford used in his gold-foil experiment? They came from a radioactive substance. Well, there are other substances that emit other kinds of particles. Some radioactive substances emit negatively-charged particles called **beta** (bay' tuh) **particles**. Others emit particles with no charge at all, which are called **gamma** (ga' muh) **particles**. Some emit two of these kinds of particles, and some emit all three.

Rutherford himself showed that gamma particles were a form of light, because they reflect off flat surfaces in the same way that light reflects off flat surfaces. Now, of course, gamma particles are not a kind of light that we can see, because they are much more energetic than the light that our eyes are designed to detect. But there are many forms of light that cannot be seen. You might have learned about infrared light, which is a form of light that is too low in energy for our eyes to see. Well, gamma particles are too high in energy for our eyes to see, but they are light, nevertheless. Because Rutherford showed that these "particles" are actually a form of light, they are now often called **gamma rays**.

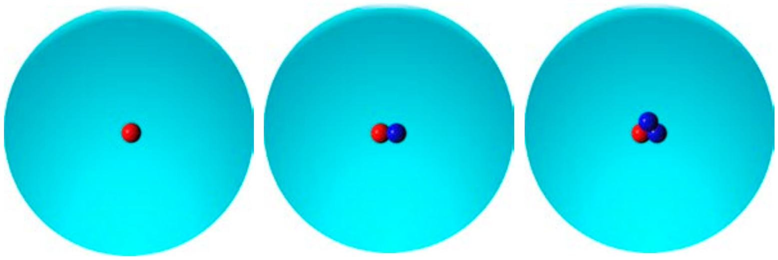
Lots of scientists were experimenting with what alpha, beta, and gamma particles could do once they were emitted by a radioactive substance. Indeed, as you already learned, Rutherford found that alpha particles could knock protons out of the nucleus of other atoms. In 1930, two German scientists, **Walther Bothe** (boh' tuh) and **Herbert Becker**, found that when they shot alpha particles into an element called beryllium (bu ril' ee uhm), a new kind of particle was produced. It had no charge, so Bothe and Becker thought that they had discovered a new kind of gamma ray.

The problem with that idea was shown by an experiment done by Pierre and Marie Curie's daughter and her husband. You might have learned about Pierre and Marie Curie in a previous course. They were among the scientists who first began to work with radioactivity. They are one of only two husband/wife couples who have won the Nobel Prize. Do you know the other husband/wife couple to have won the Nobel Prize? Their daughter, **Irène Joliot** (joh lee ette')**-Curie**, and her husband, **Frédéric Joliot**. This husband/wife duo showed that when these new uncharged particles were shot at certain substances, they could knock protons out of the nucleus.

When English scientist **James Chadwick** (who was working in Rutherford's lab) heard this, he knew that there was no way these new uncharged particles could be light. After all, light has no mass; it is pure energy. It's hard to understand how something with no mass could knock protons out of a substance. He repeated the Joliot-Curie experiments and then started seeing what other substances these uncharged particles could knock protons out of. He ended up showing mathematically that there was no way light could be knocking protons out of those substances. Instead, his results were best understood if the uncharged particles had about the same mass as a proton. These particles were also found in the nucleus of the atom, and they became known as **neutrons** (new' trahnz).

This photograph of James Chadwick was taken about 13 years after he discovered the neutron.

With Chadwick's discovery of the neutron, the idea of isotopes made perfect sense. Since the nucleus of an atom has both protons and neutrons, you could have atoms with the same number of protons, but different numbers of neutrons. The number of protons would determine the element, and the different numbers of neutrons would cause the atoms to have different masses. In Aston's experiments, for example, all the neon atoms had the same number of protons, but the heavier neon isotope had more neutrons than the lighter neon isotope.



These are the three major isotope of hydrogen. They each have one proton (red ball) in the nucleus and one electron moving in a spherical orbital around the nucleus. The lightest isotope (left) has no neutrons, the medium-mass isotope (middle) has one neutron (blue ball), and the heaviest isotope (right) has two neutrons.

So by this time in history (1932), the basic structure of the atom was finally figured out. An atom has a nucleus that contains positively-charged protons and uncharged neutrons. The number of protons in the atom determines which element the atom belongs to. All atoms with one proton, for example, are part of the element known as hydrogen. However, in addition to protons, there are usually neutrons in the nucleus. The number of neutrons in the nucleus doesn't affect which element the atom belongs to, but it does affect the mass. An atom that has just one proton and no neutrons in its nucleus, for example, is the lightest form of hydrogen. The proton gives it mass, but without any neutrons, the mass is not very large. If it has one proton and one neutron in its nucleus, it is still hydrogen, but it is a heavier form of hydrogen, because the neutron's mass adds to the proton's mass to make the atom heavier. There are also hydrogen atoms that have one proton and two neutrons in the nucleus. Those are even heavier forms of hydrogen.

In addition, the atom has negatively-charged electrons. There are always the same number of electrons as protons, so that an atom has no overall charge. The negative charge of each electron cancels out the positive charge of each proton. Those electrons travel around the nucleus, but not in fixed orbits. Instead, they travel around the nucleus in clouds that are called **orbitals**. Depending on the energy of the electrons, those orbitals can have different shapes and different sizes.

While we are still refining our understanding of how the protons and neutrons work together to make the nucleus, this is still the view of atoms that we have today. As you will see in the next chapter, this view of the atom helps to explain all sorts of things about how elements behave and how they join together to form other substances.

Go back and look at the drawing that started this chapter. Can you see what model it is based on? The electrons are in specific orbits around the nucleus, so it is basically the Bohr model of the atom. However, there are both protons and neutrons in the nucleus (the blue balls are neutrons, while the red balls are protons), which really weren't incorporated until the quantum-mechanical model became the accepted model of the atom. So it is a model of the atom that uses Bohr orbits but has the proper depiction of the nucleus.

Comprehension Check

Consider the following three atoms:

Atom A has six protons and six neutrons in its nucleus
Atom B has ten electrons traveling around a nucleus that contains 12 neutrons
Atom C has six protons and eight neutrons in its nucleus

2.14 How many protons are in the nucleus of atom B?

2.15 Which two atoms are isotopes? Which is the heavier one?

2.16 Which atom is the heaviest of them all?

Answers to the Comprehension Check Questions

2.1 Yes. It was positively charged. Your hair is made up of atoms, which have equal amounts of positive and negative charges in them. After the balloon took negative charges from your hair, there were fewer negative charges than positive charges in your hair, so your hair had an overall positive charge.

2.2 No, it would not have been attracted to the balloon. Remember, the negative charges had to move away from the balloon so that the part of the can closest to the balloon would have an overall positive charge. If the negative charges couldn't move, there would be no way for that to happen.

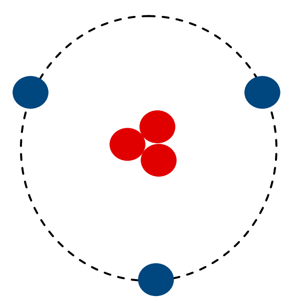
2.3 Object C weighs the least. While not the same, mass and weight are related. So the more mass there is, the more weight there is. A has more mass than B, and B has more mass than C. So C has the least mass and therefore weighs the least.

2.4 Object B floats, and object A sinks. Remember, something floats in water if it is less dense than water. It sinks if it is more dense than water.

2.5 Object B has the higher charge-to-mass ratio. If they each have the same charge, then when calculating the charge-to-mass ratio, you are taking the same number and dividing it by different masses. Since object A has a higher mass, you are dividing by a larger number. Dividing by a larger number makes a smaller number. Similarly, dividing by a smaller number makes a large number.

2.6 The charge on the object is 15 times *e*. Remember, Millikan showed that you give objects a negative charge by adding electrons to it. The object gains a charge of *e* for every electron it is given. If you put 15 electrons on the object, you are giving it 15 times the charge of an electron.

2.7 The atom would have a charge. It would be positive. An atom starts out with equal numbers of protons and electrons. Since the charges of the electrons and protons cancel one another out, there is no overall charge. If you remove an electron, then there is one proton that has no electron to cancel it out. That results in an overall positive charge.



2.8 Remember, atoms have the same number of protons as electrons. That way, they have no overall charge. So if it has three electrons, it also has three protons in its nucleus. The three electrons orbit the nucleus, so it would look something like the picture on the right. Please note that this is not the correct model of the atom. You will learn that model soon.

2.9 It would gain energy. Remember, light is a form of energy, so if the particle absorbs light, it absorbs energy, which would add to the energy it already has. When particles emit light, they lose energy, and when they absorb light, they gain energy.

2.10 It absorbed energy. The farther from the nucleus, the higher the energy of the orbit. So orbits close to the nucleus are lower in energy, and orbits far from the nucleus are higher in energy. To go from an orbit close to the nucleus to one far from the nucleus requires gaining energy.

2.11 It is moving closer to the nucleus. If an electron is emitting light, it is losing energy. That means it is moving from a higher-energy orbit far from the nucleus to a lower-energy orbit closer to the nucleus.

2.12 They should happen at multiples of a different number of volts. Remember, each atom has its own set of energies of light that it releases. That means each atom has its own set of energies for the electrons' orbits. Thus, the energy that allows a neon electron to move to another orbit will be different from the energy that allows a mercury electron to move to another orbit. As a result, neon will absorb different amounts of energy than mercury.

2.13 The electron in the larger s orbital will have more energy. Even though this isn't the Bohr model, Bohr was right that the farther from the nucleus, the more energy the electron must have. Since the electron in the larger s orbital can be farther from the nucleus, it must have more energy.

2.14 There are ten protons. Atoms must have the same number of protons as electrons so that the overall charge is zero. Since there are ten electrons, there must be ten protons.

2.15 A and C are isotopes, and C is the heavier one. Isotopes are atoms that belong to the same element but have different masses. An element is defined by the number of protons, so A and C are in the same element, since they each have six protons. The masses are different, however, since C has two more neutrons than A. That makes C heavier than A.

2.16 B is the heaviest of them all. Each proton adds mass to the atom, each neutron adds mass, and each electron adds mass. So the atom that has the most protons, neutrons, and electrons is the heaviest. B has four more protons and four more electrons than A and C. It also has six more neutrons than A and four more than C.

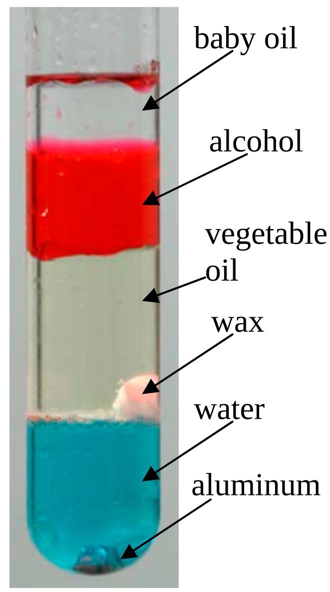
Chapter Review

1. Define the following terms:

a. Model	d. Mass	g. Nucleus	j. Isotopes
b. Matter	e. Ratio	h. Centripetal force	
c. Conductor	f. Density	i. Mass spectrograph	

2. Dalton's atomic theory was important, but it had two serious errors. What were those two errors?

3. If you remove electrons from something that has no charge, will it develop a charge? If so, what kind of charge (positive or negative) will it develop?



4. Object A weighs significantly more than object B. Which has less mass?

5. True or False: Mass and weight are the same.

6. You see a glass tube with solids and liquids in it, as shown on the right. Order all the substances in the glass according to their densities. Start with the substance that has the lowest density and finish with the substance that has the highest density.

7. Looking at the glass tube on the right again, suppose you have equal volumes of wax and alcohol. Which has more mass?

8. When you are making an object negatively charged, what are you adding to the object?

9. Can an object have any amount of negative charge you want to give it?

10. What kind of radioactive particles did Rutherford's gold-foil experiment use? Were they positive, negative, or not charged?

11. To come up with his model of the atom, what did Bohr assume about where the electrons could orbit the nucleus?

12. In the Bohr model of the atom, an electron in the first orbit (the one closest to the nucleus) has an energy of E_1. In the second orbit (which is farther from the nucleus), it has an energy of E_2.

   a. Which is larger: E_1 or E_2?
   b. If an electron wants to move from the first Bohr orbit to the second, does it lose energy or gain energy?

13. An electron is emitting light. Is it moving towards the nucleus or away from the nucleus?

14. What did the Franck-Hertz experiment demonstrate about mercury atoms?

15. How did Schrödinger treat electrons in order to produce the quantum-mechanical model of the atom?

16. When waves overlap in an interesting way, what do we call the pattern they form?

17. What do we usually represent the quantum-mechanical model's electron clouds with?

18. What particles are found in the nucleus of the atom? What are their charges?

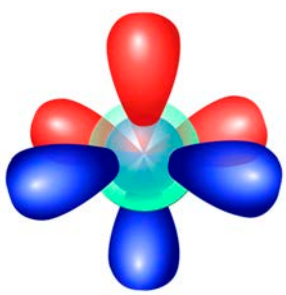
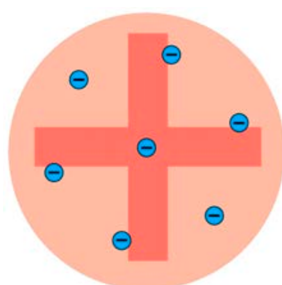
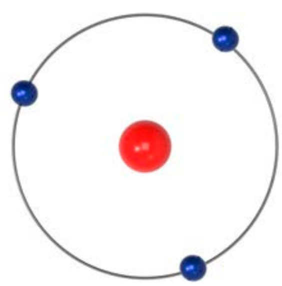
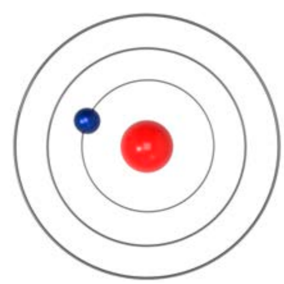
19. Atom A has eight electrons. There are also 8 neutrons in its nucleus.

    a. How many protons are in its nucleus?
    b. Another atom has 7 protons and 9 neutrons in its nucleus. Is it an isotope of atom A?
    c. Another atom has 8 protons and 10 neutrons in its nucleus. Is it an isotope of atom A?

20. Atom A and B are isotopes. A is heavier than B.

    a. Compare the number of protons in each atom.
    b. Compare the number of electrons in each atom.
    c. Compare the number of neutrons in each atom.

21. Four illustrations of different models for the atom are shown below. Put them in chronological order, starting with the oldest one. Also, name them.

a. 
b. 
c. 
d. 

I just want to remind you that if you didn't understand everything you read in the chapter, that's fine. This review is the key to what I want you to understand. So if you are comfortable with the material covered in this review, you are ready for the test.