CK-12 Foundation

The Atomic Theory

Say Thanks to the Authors
Click http://www.ck12.org/saythanks
(No sign in required)

Bewick **Parsons**

To access a customizable version of this book, as well as other interactive content, visit www.ck12.org

Printed: June 24, 2012

Authors
Sharon Bewick, Richard Parsons

Editor
Shonna Robinson

Contents

Chapter 1

The Atomic Theory

1.1 The Atomic Theory

Lesson Objectives

The student will:

- give a short history of how the concept of the atom developed.
- describe the contributions of Democritus and Dalton to the atomic theory.
- summarize Dalton's atomic theory and explain its historical development.
- state the law of definite proportions.
- state the law of multiple proportions.

Vocabulary

- atomos
- Dalton's atomic theory
- law of definite proportions
- law of multiple proportions

Introduction

You learned earlier in the chapter "Matter and Energy" that all matter in the universe is made up of tiny building blocks called atoms. All modern scientists accept the concept of the atom, but when the concept of the atom was first proposed about 2,500 years ago, ancient philosophers laughed at the idea. After all, it is difficult to be convinced that something too small to be seen really exists. We will spend some time considering the evidence (observations) that convinced scientists of the existence of atoms.

Democritus and the Greek Philosophers

Before we discuss the experiments and evidence that have convinced scientists matter is made up of atoms, it is only fair to credit the man who proposed the concept of the atom in the first place. About 2,500 years ago, early Greek philosophers believed the entire universe was a single, huge entity. In other words,

“everything was one.” They believed that all objects, all matter, and all substances were connected as a single, big, unchangeable “thing.”

Figure 1.1: Democritus was known as

One of the first people to propose the existence of atoms was a man known as Democritus, pictured in **Figure** 1.1. He suggested an alternative theory where **atomos** – tiny, indivisible, solid objects – made up all matter in the universe. Democritus then reasoned that changes occur when the many atomos in an object were reconnected or recombined in different ways. Democritus even extended his theory to suggest that there were different varieties of atomos with different shapes, sizes, and masses. He thought, however, that shape, size, and mass were the only properties differentiating the types of atomos. According to Democritus, other characteristics, like color and taste, did not reflect properties of the atomos themselves but from the different ways in which the atomos were combined and connected to one another.

So how could the Greek philosophers have known that Democritus had a good idea with his theory of atomos? The best way would have been to take some careful observation and conduct a few experiments. Recall, however, that the early Greek philosophers tried to understand the nature of the world through reason and logic, not through experimentation and observation. The Greek philosophers truly believed that, above all else, our understanding of the world should rely on logic. In fact, they argued that the world couldn't be understood using our senses at all because our senses could deceive us. Therefore, instead of relying on observation, Greek philosophers tried to understand the world using their minds and, more specifically, the power of reason (see **Figure** 1.2).

As a result, the early Greek philosophers developed some very interesting ideas, but they felt no need to justify their ideas. You may recall from the “Introduction to Chemistry” chapter that Aristotle concluded men had more teeth than women did. He concluded this without ever checking in anyone's mouth because his conclusion was the “logical” one. As a result, the Greek philosophers missed or rejected a lot of discoveries that could have made otherwise because they never performed any experiments. Democritus's theory would be one of these rejected theories. It would take over two millennia before the theory of atomos (or atoms, as they're known today) was fully appreciated.

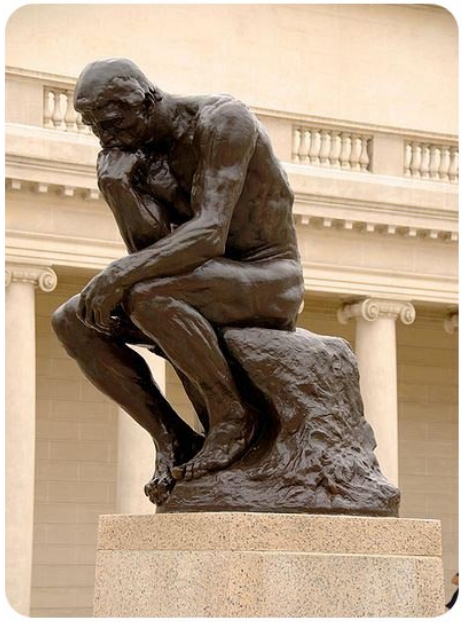


Figure 1.2: This sculpture (named

Dalton's Atomic Theory

Let's consider a simple but important experiment that suggested matter might be made up of atoms. In the late 1700s and early 1800s, scientists began noticing that when certain substances, like hydrogen and oxygen, were combined to produce a new substance, the reactants (hydrogen and oxygen) always reacted in the same proportions by mass. In other words, if 1 gram of hydrogen reacted with 8 grams of oxygen, then 2 grams of hydrogen would react with 16 grams of oxygen, and 3 grams of hydrogen would react with 24 grams of oxygen.

Strangely, the observation that hydrogen and oxygen always reacted in the "same proportions by mass" wasn't unique to hydrogen and oxygen. In fact, it turned out that the reactants in every chemical reaction for a given compound react in the same proportions by mass. Take, for example, nitrogen and hydrogen, which can react to produce ammonia (NH_3). In chemical reactions, 1 gram of hydrogen will react with 4.7 grams of nitrogen, and 2 grams of hydrogen will react with 9.4 grams of nitrogen. Can you guess how much nitrogen would react with 3 grams of hydrogen?

Scientists studied reaction after reaction, but every time the result was the same. The reactants always reacted in the same proportions by mass or in what we call "definite proportions," as illustrated in **Figure** 1.3. As a result, scientists proposed the **law of definite proportions**. This law states that:

In a given type of chemical substance, the elements always combine in the same proportions by mass.

This version of the law is a more modern version. Earlier, you learned that an element is a substance made up of only one type of atom, but when the law of definite proportions was first discovered, scientists did not know about atoms or elements and stated the law slightly differently. We'll stick with this modern version, though, since it is the easiest version to understand.

The law of definite proportions applies when the elements reacting together form the same product. Therefore, the law of definite proportions can be used to compare two experiments in which hydrogen and oxygen react to form water. The law, however, cannot be used to compare one experiment in which hydrogen and oxygen react to form water with another experiment in which hydrogen and oxygen react to form hydrogen

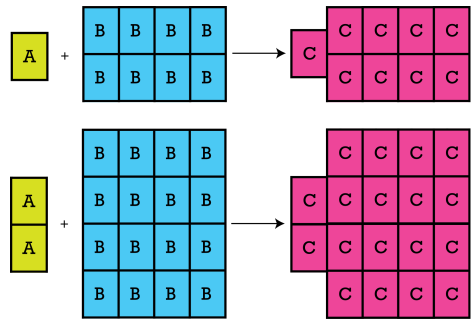


Figure 1.3: If 1 gram of A reacts with 8 grams of B, then by the law of definite proportions, 2 grams of A must react with 16 grams of B. If 1 gram of A reacts with 8 grams of B, then by the law of conservation of mass, they must produce 9 grams of C.

peroxide (peroxide is another material that can be made from hydrogen and oxygen).

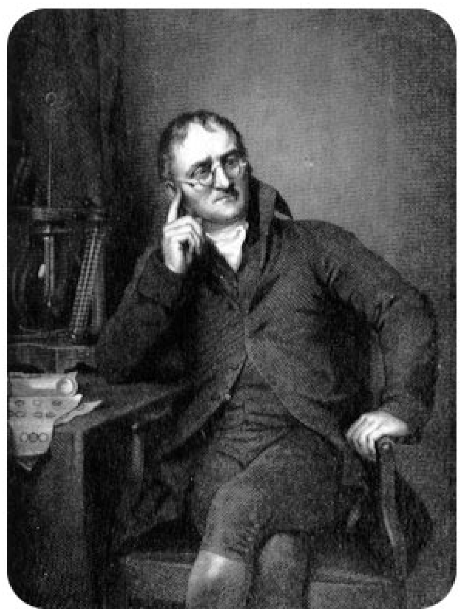


Figure 1.4: Unlike the early Greek philosophers, John Dalton was both a thinker and an experimenter. He would help develop the modern conception of an atom based on his experimental results.

A man named John Dalton (**Figure** 1.4) discovered this limitation in the law of definite proportions in some of his experiments. Dalton was experimenting with several reactions in which the reactant elements formed different products, depending on the experimental conditions he used. One common reaction that he studied was the reaction between carbon and oxygen. When carbon and oxygen react, they produce two different substances – we'll call these substances *A* and *B*. It turned out that, given the same amount of carbon, forming *B* always required exactly twice as much oxygen as forming *A*. In other words, if you could make *A* with 3 grams of carbon and 4 grams of oxygen, *B* could be made with the same 3 grams of carbon but with 8 grams of oxygen instead. Dalton asked himself – why does *B* require twice as much oxygen as *A* does? Why not 1.21 times as much oxygen, or 0.95 times as much oxygen? Why a whole number like 2?

The situation became even stranger when Dalton tried similar experiments with different substances. For example, when he reacted nitrogen and oxygen, Dalton discovered that he could make three different

substances – we'll call them *C*, *D*, and *E*. As it turned out, for the same amount of nitrogen, *D* always required twice as much oxygen as C does. Similarly, *E* always required exactly four times as much oxygen as *C* does. Once again, Dalton noticed that small whole numbers (2 and 4) seemed to be the rule. Dalton used his experimental results to propose the **law of multiple proportions**:

When two elements react to form more than one substance and the same amount of one element (like oxygen) is used in each substance, then the ratio of the masses used of the other element (like nitrogen) will be in small whole numbers.

This law summarized Dalton's findings, but it did not explain why the ratio was a small whole number. Dalton thought about his law of multiple proportions and tried to develop a theory that would explain it. Dalton also knew about the law of definite proportions and the law of conservation of mass, so what he really wanted was a theory that explained all three laws with a simple, plausible model. One way to explain the relationships that Dalton and others had observed was to suggest that materials like nitrogen, carbon, and oxygen were composed of small, indivisible quantities, which Dalton called "atoms" (in reference to Democritus's original idea). Dalton used this idea to generate what is now known as **Dalton's atomic theory**.

Dalton's atomic theory:

1. Matter is made of tiny particles called atoms.
2. Atoms are indivisible. During a chemical reaction, atoms are rearranged, but they do not break apart, nor are they created or destroyed.
3. All atoms of a given element are identical in mass and other properties.
4. The atoms of different elements differ in mass and other properties.
5. Atoms of one element can combine with atoms of another element to form compounds. In a given compound, however, the different types of atoms are always present in the same relative numbers.

Historical note: Some people think that Dalton developed his atomic theory before stating the law of multiple proportions, while others argue that the law of multiple proportions, though not formally stated, was actually discovered first. In reality, Dalton was probably contemplating both concepts at the same time, although it is hard to say conclusively from looking at the laboratory notes he left behind.

Lesson Summary

- 2,500 years ago, Democritus suggested that all matter in the universe was made up of tiny, indivisible, solid objects he called atomos.
- Other Greek philosophers disliked Democritus's atomos theory because they felt it was illogical.
- The law of definite proportions states that in a given chemical substance, the elements are always combined in the same proportions by mass.
- The law of multiple proportions states that when two elements react to form more than one substance and the same amount of one element is used in each substance, then the ratio of the masses used of the other element will be in small whole numbers.
- Dalton used the law of definite proportions, the law of multiple proportions, and the law of conservation of mass to propose his atomic theory.
- Dalton's atomic theory states:
    1. Matter is made of tiny particles called atoms.
    2. Atoms are indivisible. During a chemical reaction, atoms are rearranged, but they do not break apart, nor are they created or destroyed.
    3. All atoms of a given element are identical in mass and other properties.
    4. The atoms of different elements differ in mass and other properties.

5. Atoms of one element can combine with atoms of another element to form compounds. In a given compound, however, the different types of atoms are always present in the same relative numbers.

Review Questions

1. It turns out that a few of the ideas in Dalton's atomic theory aren't entirely correct. Are inaccurate theories an indication that science is a waste of time?
2. Suppose you are trying to decide whether to wear a sweater or a T-shirt. To make your decision, you phone two friends. The first friend says, "Wear a sweater, because I've already been outside today, and it's cold." The second friend, however, says, "Wear a T-shirt. It isn't logical to wear a sweater in July." Would you decide to go with your first friend and wear a sweater or with your second friend and wear a T-shirt? Why?
3. Decide whether each of the following statements is true or false.
   (a) Democritus believed that all matter was made of atomos.
   (b) Democritus also believed that there was only one kind of atomos.
   (c) Most early Greek scholars thought that the world was "ever-changing."
   (d) If the early Greek philosophers hadn't been so interested in making gold, they probably would have liked the idea of the atomos.
4. Match the person, or group of people, with their role in the development of chemistry in the table below.

Person/Group of People	**Role in Chemistry**
(a) early Greek philosophers	(i) first suggested that all matter was made up of tiny, indivisible, solid objects
(b) alchemists	(ii) tried to apply logic to the world around them
(c) John Dalton	(iii) proposed the first scientific theory relating chemical changes to the structure, properties, and behavior of atoms
(d) Democritus	(iv) were primarily concerned with finding ways to turn common metals into gold

5. Early Greek philosophers felt that Democritus's atomos theory was illogical because:
   (a) no matter how hard they tried, they could never break matter into smaller pieces.
   (b) it didn't help them to make gold.
   (c) sulfur is yellow and carbon is black, so clearly atomos must be colored.
   (d) empty space is illogical because it implies that nothing is actually something.
6. Identify the law that explains the following observation: Carbon monoxide can be formed by reacting 12 grams of carbon with 16 grams of oxygen. To form carbon dioxide, however, 12 grams of carbon must react with 32 grams of oxygen.
7. Identify the law that explains the following observation: Carbon monoxide can be formed by reacting 12 grams of carbon with 16 grams of oxygen. It can also be formed by reacting 24 grams of carbon with 32 grams of oxygen.
8. Identify the law that explains the following observation: 28 grams of carbon monoxide are formed when 12 grams of carbon reacts with 16 grams of oxygen.

9. Identify the law that explains the following observations: When 12 grams of carbon react with 4 grams of hydrogen, they produce methane, and there is no carbon or hydrogen left over at the end of the reaction. If, however, 11 grams of carbon react with 4 grams of hydrogen, there is hydrogen left over at the end of the reaction.
10. Which of the following is *not* part of Dalton's atomic theory?
    (a) Matter is made of tiny particles called atoms.
    (b) During a chemical reaction, atoms are rearranged.
    (c) During a nuclear reaction, atoms are split apart.
    (d) All atoms of a specific element are the same.
11. Consider the following data: 3.6 grams of boron react with 1.0 grams of hydrogen to give 4.6 grams of BH_3. How many grams of boron would react with 2.0 grams of hydrogen?
12. Consider the following data: 12 grams of carbon and 4 grams of hydrogen react to give 16 grams of compound A. 24 grams of carbon and 6 grams of hydrogen react to give 30 grams of compound B. Are compound A and compound B the same? Why or why not?

1.2 Further Understanding of the Atom

Lesson Objectives

The student will:

- explain the observations that led to Thomson's discovery of the electron.
- describe Thomson's plum-pudding model of the atom.
- draw a diagram of Thomson's plum-pudding model of the atom and explain why it has this name.
- describe Rutherford's gold foil experiment and explain how this experiment disproved the plum-pudding model.
- draw a diagram of the Rutherford model of the atom and label the nucleus and the electron cloud.

Vocabulary

- cathode ray tube
- electron
- nucleus
- proton
- subatomic particle

Introduction

Dalton's atomic theory held up well in a lot of the different chemical experiments that scientists performed to test it. For almost 100 years, it seemed as if Dalton's atomic theory was the whole truth. It wasn't until 1897 when a scientist named J. J. Thomson conducted some research that suggested Dalton's atomic theory wasn't the entire story. Dalton had gotten a lot right - he was right in saying matter is made up of atoms; he was right in saying there are different kinds of atoms with different mass and other properties; he was *almost* right in saying atoms of a given element are identical; he was right in saying that atoms are merely rearranged during a chemical reaction; and he was right in saying a given compound always has atoms present in the same relative numbers. But he was *wrong* in saying atoms were indivisible or indestructible. As it turns out, atoms are divisible. In fact, atoms are composed of even smaller, more

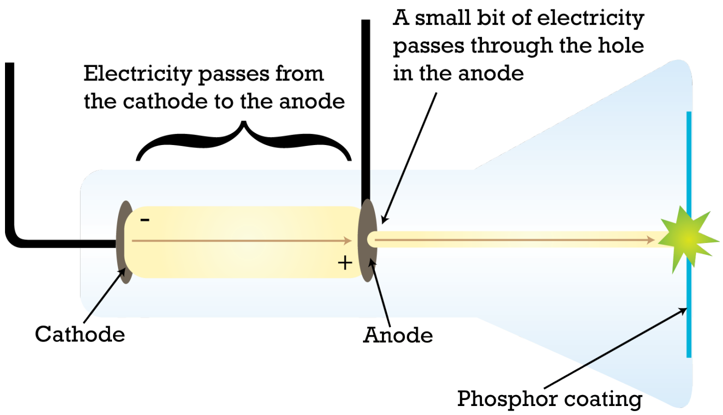
fundamental particles. These particles, called **subatomic particles**, are particles that are smaller than the atom. The discoveries of these subatomic particles are the focus of this chapter.

Thomson's Plum-Pudding Model

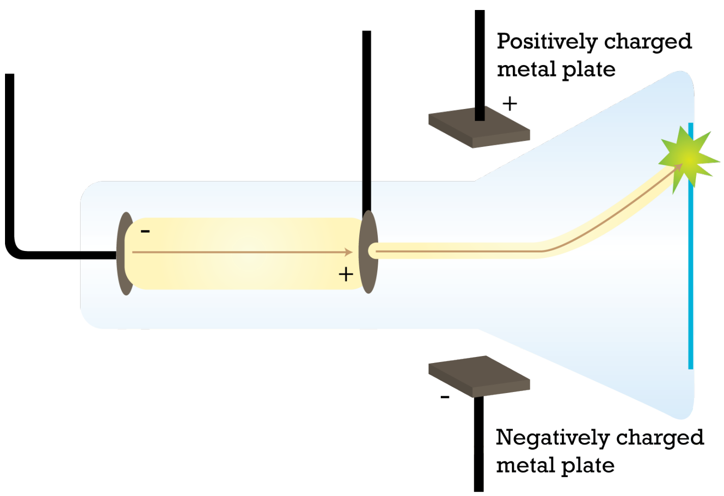
In the mid-1800s, scientists were beginning to realize that the study of chemistry and the study of electricity were actually related. First, a man named Michael Faraday showed how passing electricity through mixtures of different chemicals could cause chemical reactions. Shortly after that, scientists found that by forcing electricity through a tube filled with gas, the electricity made the gas glow. Scientists didn't, however, understand the relationship between chemicals and electricity until a British physicist named J. J. Thomson began experimenting with what is known as a cathode ray tube (**Figure** 1.5).

Figure 1.5: A portrait of J. J. Thomson.

The figure below shows a basic diagram of a cathode ray tube like the one Thomson would have used. A **cathode ray tube** is a small glass tube with a cathode (a negatively charged metal plate) and an anode (a positively charged metal plate) at opposite ends. By separating the cathode and anode a short distance, the cathode ray tube can generate what are known as cathode rays – rays of electricity that flow from the cathode to the anode. Thomson wanted to know what cathode rays were, where cathode rays came from, and whether cathode rays had any mass or charge. The techniques that he used to answer these questions were very clever and earned him a Nobel Prize in physics. First, by cutting a small hole in the anode, Thomson found that he could get some of the cathode rays to flow through the hole in the anode and into the other end of the glass cathode ray tube. Next, he figured out that if he painted a substance known as phosphor onto the far end of the cathode ray tube, he could see exactly where the cathode rays hit because the cathode rays made the phosphor glow.



Thomson must have suspected that cathode rays were charged, because his next step was to place a positively charged metal plate on one side of the cathode ray tube and a negatively charged metal plate on the other side, as shown below. The metal plates didn't actually touch the cathode ray tube, but they were close enough that a remarkable thing happened. The flow of the cathode rays passing through the hole in the anode was bent upwards towards the positive metal plate and away from the negative metal plate. In other words, instead of glowing directly across from the hole in the anode, the phosphor now glowed at a spot quite a bit higher in the tube.



Thomson thought about his results for a long time. It was almost as if the cathode rays were attracted to the positively charged metal plate and repelled from the negatively charged metal plate. Thomson knew that charged objects are attracted to and repelled from other charged objects according to the rule: opposite charges attract, like charges repel. This means that a positive charge is attracted to a negative charge but repelled from another positive charge. Similarly, a negative charge is attracted to a positive charge but repelled from another negative charge. Using the "opposite charges attract, like charges repel" rule, Thomson argued that if the cathode rays were attracted to the positively charged metal plate and repelled from the negatively charged metal plate, the rays themselves must have a negative charge.

Thomson then did some rather complex experiments with magnets and used the results to prove that cathode rays not only were negatively charged, but they also had mass. Remember that anything with mass is part of what we call matter. In other words, these cathode rays must be the result of negatively charged matter flowing from the cathode to the anode. It was here that Thomson encountered a problem. According to his measurements, these cathode rays either had a ridiculously high charge or very, very little mass – much less mass than the smallest known atom. How was this possible? How could the matter making up cathode rays be smaller than an atom if atoms were indivisible? Thomson made a radical

proposal: maybe atoms are divisible. He suggested that the small, negatively charged particles making up the cathode ray were actually pieces of atoms. He called these pieces "corpuscles," although today we know them as **electrons**. Thanks to his clever experiments and careful reasoning, Thomson is credited with the discovery of the electron.

For a demonstration of cathode ray tubes **(1h)**, see http://www.youtube.com/watch?v=XU8nMKkzbT8 (1:09).

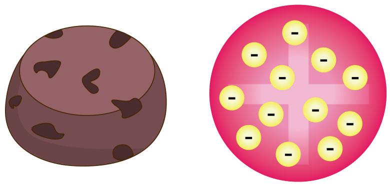
Figure 1.6: (Watch Youtube Video)
http://www.ck12.org/flexbook/embed/view/355

Now imagine what would happen if atoms were made entirely of electrons. First of all, electrons are very, very small; in fact, electrons are about 2,000 times smaller than the smallest known atom, so every atom would have to contain a lot of electrons. But there's another, bigger problem: electrons are negatively charged. Therefore, if atoms were made entirely out of electrons, the atoms themselves would be negatively charged, which would mean all matter was negatively charged as well.

Of course, matter isn't negatively charged. Most matter is what we call neutral – it has no charge at all. How can matter be neutral if matter is composed of atoms and atoms are composed of negative electrons? The only possible explanation is that atoms must consist of more than just electrons. Atoms must also contain some type of positively charged material that balances the negative charge of the electrons. Negative and positive charges of equal size cancel each other out, just like negative and positive numbers of equal size. If an atom contains an electron with a -1 charge and some form of material with a $+1$ charge, overall the atom must have a $(+1) + (-1) = 0$ charge. In other words, the atom would be neutral, or have no overall charge.

Based on the fact that atoms are neutral and based on Thomson's discovery that atoms contain negative subatomic particles called electrons, scientists assumed that atoms must also contain a positive substance. It turned out that this positive substance was another kind of subatomic particle known as the **proton**. Although scientists knew that atoms had to contain positive material, protons weren't actually discovered, or understood, until quite a bit later.

When Thomson discovered the negative electron, he also realized that atoms had to contain positive material as well. As a result, Thomson formulated what's known as the plum-pudding model for the atom. According to the plum-pudding model, the negative electrons were like pieces of fruit and the positive material was like the batter or the pudding. In the figure below, an illustration of a plum pudding is on the left and an illustration of Thomson's plum-pudding model is on the right. (Instead of a plum pudding, you can also think of a chocolate chip cookie. In that case, the positive material in the atom would be the batter in the chocolate chip cookie, while the negative electrons would be scattered through the batter like chocolate chips.)



This made a lot of sense given Thomson's experiments and observations. Thomson had been able to isolate electrons using a cathode ray tube; however, he had never managed to isolate positive particles. Notice in the image above how easy it would be to pick the pieces of fruit out of a plum pudding. On the other hand, it would be a lot harder to pick the batter out of the plum pudding because the batter is everywhere. If an atom were similar to a plum pudding in which the electrons are scattered throughout the "batter" of positive material, then you would expect it to be easy to pick out the electrons and a lot harder to pick out the positive material.

Everything about Thomson's experiments suggested the plum-pudding model was correct. According to the scientific method, however, any new theory or model should be tested by further experimentation and observation. In the case of the plum-pudding model, it would take a man named Ernest Rutherford to prove it wrong. Rutherford and his experiments will be the topic of the next section.

There was one thing that Thomson was unable to determine. He had measured the charge-to-mass ratio of the electron, but he had been unable to measure accurately the charge on the electron. Instead, a different scientist named Robert Millikan would determine the charge of the electron with his oil drop experiment. When combined with Thomson's charge-to-mass ratio, Millikan was able to calculate the mass of the electron. Millikan's experiment involved putting charges on tiny droplets of oil suspended between charged metal plates and measuring their rate of descent. By varying the charge on different drops, he noticed that the electric charges on the drops were all multiples of 1.6×10^{-19} C (coulomb), the charge of a single electron.

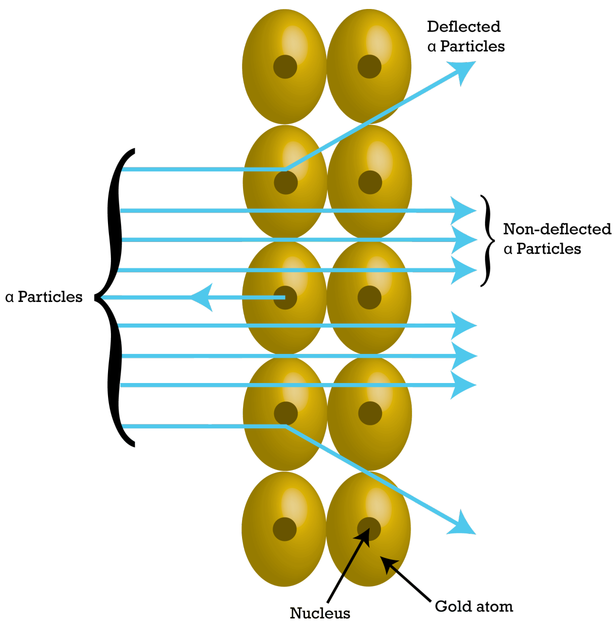
Rutherford's Nuclear Model

Disproving Thomson's plum-pudding model began with the discovery that an element known as uranium emits positively charged particles called alpha particles as it undergoes radioactive decay. Radioactive decay occurs when one element decomposes into another element. It only happens with a few very unstable elements. Alpha particles themselves didn't prove anything about the structure of the atom, but they were used to conduct some very interesting experiments.

Ernest Rutherford (pictured in **Figure** 1.7) was fascinated by all aspects of alpha particles and used them as tiny bullets that could be fired at all kinds of different materials. The results of one experiment in particular surprised Rutherford and everyone else.

Rutherford found that when he fired alpha particles at a very thin piece of gold foil, an interesting phenomenon happened. The diagram below helps illustrate Rutherford's findings. Almost all of the alpha particles went straight through the foil as if they had hit nothing at all. Every so often, though, one of the alpha particles would be deflected slightly as if it had bounced off something hard. Even less often, Rutherford observed alpha particles bouncing straight back at the "gun" from which they had been fired. It was as if these alpha particles had hit a wall head-on and had ricocheted right back in the direction that they had come from.

Figure 1.7: A portrait of Ernest Rutherford.

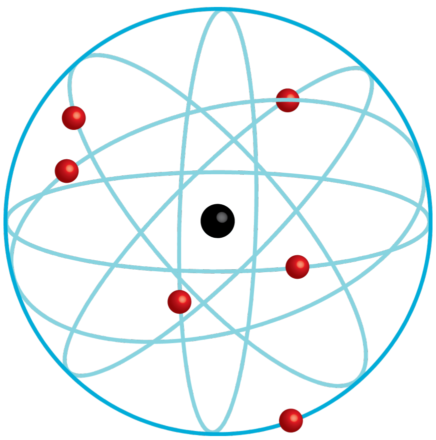


Rutherford thought that these experimental results were rather odd. He expected firing alpha particles at gold foil to be like shooting a high-powered rifle at tissue paper. The bullets would break through the tissue paper and keep on going, almost as if they had hit nothing at all. That was what Rutherford had expected to see when he fired alpha particles at the gold foil. The fact that most alpha particles passed through did not shock him, but how could he explain the alpha particles that were deflected? Furthermore, how could he explain the alpha particles that bounced right back as if they had hit a wall?

Rutherford decided that the only way to explain his results was to assume that the positive matter forming the gold atoms was not distributed like the batter in plum pudding. Instead, he proposed that the positive matter was concentrated in one spot, forming a small, positively charged particle somewhere in the center

of the gold atom. We now call this clump of positively charged mass the **nucleus**. According to Rutherford, the presence of a nucleus explained his experiments because it implied that most of the positively charged alpha particles would pass through the gold foil without hitting anything at all. Occasionally, though, the alpha particles would actually collide with a gold nucleus, causing the alpha particles to be deflected or even bounced back in the direction they came from.

While Rutherford's discovery of the positively charged atomic nucleus offered insight into the structure of the atom, it also led to some questions. According to the plum-pudding model, electrons were like plums embedded in the positive batter of the atom. Rutherford's model, though, suggested that the positive charge was concentrated into a tiny particle at the center of the atom, while most of the rest of the atom was empty space. What did that mean for the electrons? If they weren't embedded in the positive material, exactly what were they doing? How were they held in the atom? Rutherford suggested that the electrons might be circling or orbiting the positively charged nucleus as some type of negatively charged cloud, like in the image below. At the time, however, there wasn't much evidence to suggest exactly how the electrons were held in the atom.



A short animation of Rutherford's experiment **(1h)** can be found at http://www.youtube.com/watch?v=5pZj0u_-XMbc (0:47).

Figure 1.8: (Watch Youtube Video)
http://www.ck12.org/flexbook/embed/view/357

For another video discussing J.J. Thomson's use of a cathode ray tube in his discovery of the electron **(1h)**, see http://www.youtube.com/watch?v=IdTxGJjA4Jw (2:54).

Despite the problems and questions associated with Rutherford's experiments, his work with alpha particles seemed to point to the existence of an atomic nucleus. Between Thomson, who discovered the electron, and Rutherford, who suggested that the positive charges were concentrated at the atom's center, the 1890s and early 1900s saw huge steps in understanding the atom at the subatomic (smaller than the size of an atom) level. Although there was still some uncertainty with respect to exactly how subatomic particles were organized in the atom, it was becoming more and more obvious that atoms were indeed divisible. Moreover,

Figure 1.9: (Watch Youtube Video)
http://www.ck12.org/flexbook/embed/view/358

it was clear that an atom contained negatively charged electrons and a positively charged nucleus. In the next lesson, we'll look more carefully at the structure of the nucleus. We'll learn that while the atom is made up of positive and negative particles, it also contains neutral particles that neither Thomson nor Rutherford were able to detect with their experiments.

Lesson Summary

- Dalton's atomic theory wasn't entirely correct, as it was found that atoms can be divided into smaller subatomic particles.
- According to Thomson's plum-pudding model, the negatively charged electrons in an atom are like the pieces of fruit in a plum pudding, while the positively charged material is like the batter.
- When Ernest Rutherford fired alpha particles at a thin gold foil, most alpha particles went straight through; however, a few were scattered at different angles, and some even bounced straight back.
- In order to explain the results of his gold foil experiment, Rutherford suggested that the positive matter in the gold atoms was concentrated at the center of the gold atom in what we now call the nucleus of the atom.

Further Reading / Supplemental Links

The *learner.org* website allows users to view the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The video called "The Atom" explores the structure of the atom.

- http://learner.org/resources/series61.html

Review Questions

1. Decide whether each of the following statements is true or false.
   (a) Cathode rays are positively charged.
   (b) Cathode rays are rays of light, thus they have no mass.
   (c) Cathode rays can be repelled by a negatively charged metal plate.
   (d) J.J. Thomson is credited with the discovery of the electron.
   (e) Phosphor is a material that glows when struck by cathode rays.
2. Match each observation with the correct conclusion.
   (a) Cathode rays are attracted to a positively charged metal plate.
      i. Cathode rays are positively charged.

ii. Cathode rays are negatively charged.
iii. Cathode rays have no charge.

(b) Electrons have a negative charge.

i. Atoms must be negatively charged.
ii. Atoms must be positively charged.
iii. Atoms must also contain positive subatomic material.

(c) Alpha particles fired at a thin gold foil are occasionally scattered back in the direction that they came from.

i. The positive material in an atom is spread throughout like the batter in pudding.
ii. Atoms contain neutrons.
iii. The positive charge in an atom is concentrated in a small area at the center of the atom.

3. What is the name given to the tiny clump of positive material at the center of an atom?
4. Choose the correct statement.

(a) Ernest Rutherford discovered the atomic nucleus by performing experiments with aluminum foil.
(b) Ernest Rutherford discovered the atomic nucleus using a cathode ray tube.
(c) When alpha particles are fired at a thin gold foil, they never go through.
(d) Ernest Rutherford proved that the plum-pudding model was incorrect.
(e) Ernest Rutherford experimented by firing cathode rays at gold foil.

5. Answer the following questions:

(a) Will the charges $+2$ and -2 cancel each other out?
(b) Will the charges $+2$ and -1 cancel each other out?
(c) Will the charges $+1$ and $+1$ cancel each other out?
(d) Will the charges -1 and $+3$ cancel each other out?
(e) Will the charges $+9$ and -9 cancel each other out?

6. Electrons are _________ negatively charged metals plates and _________ positively charged metal plates.
7. What was J. J. Thomson's name for electrons?
8. A "sodium cation" is a sodium atom that has lost one of its electrons. Would the charge on a sodium cation be positive, negative, or neutral? Would sodium cations be attracted to a negative metal plate or a positive metal plate? Would electrons be attracted to or repelled from sodium cations?
9. Suppose you have a cathode ray tube coated with phosphor so that you can see where on the tube the cathode ray hits by looking for the glowing spot. What will happen to the position of this glowing spot if:

(a) a negatively charged metal plate is placed above the cathode ray tube?
(b) a negatively charged metal plate is placed to the right of the cathode ray tube?
(c) a positively charged metal plate is placed to the right of the cathode ray tube?
(d) a negatively charged metal plate is placed above the cathode ray tube, and a positively charged metal plate is placed to the left of the cathode ray tube?
(e) a positively charged metal plate is placed below the cathode ray tube, and a positively charged metal plate is also placed to the left of the cathode ray tube?

1.3 Atomic Structure

Lesson Objectives

The student will:

- identify the three major subatomic particles and their charges, masses, and location in the atom.
- briefly describe the discovery of the neutron.
- define atomic number.
- describe the size of the nucleus in relation to the size of the atom.
- explain what is meant by the atomic mass of an element and describe how atomic masses are related to carbon-12.
- define mass number.
- explain what isotopes are and how isotopes affect an element's atomic mass.
- determine the number of protons, neutrons, and electrons in an atom.
- calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

Vocabulary

- atomic mass
- atomic mass unit
- atomic number
- dalton
- isotopes
- mass number
- neutron
- strong nuclear force

Introduction

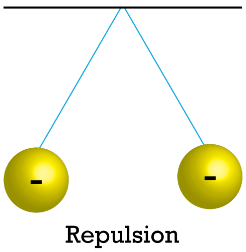
Dalton's atomic theory explained a lot about matter, chemicals, and chemical reactions. Nevertheless, it wasn't entirely accurate because, contrary to what Dalton believed, atoms can in fact be broken apart into smaller subunits or subatomic particles. The first type of subatomic particle to be found in an atom was the negatively charged electron. Since atoms are neutral, though, they must also contain positive material. In his gold foil experiment, Rutherford proved that the positive substance in an atom was concentrated in a small area at the center of the atom, leaving most the rest of the atom as empty space. In this lesson, we'll examine the subatomic particles making up the atom a little more closely.

This video gives basic information about the nucleus of atoms including comparative sizes of atom vs nucleus **(1e)**, see http://www.youtube.com/watch?v=Tfy0sIVfVOY (2:03).

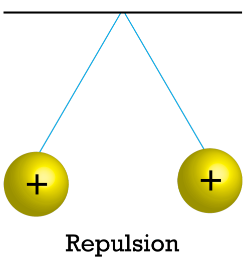
Figure 1.10: (Watch Youtube Video)
http://www.ck12.org/flexbook/embed/view/359

Electrons, Protons, and Neutrons

Electrons have a negative charge. As a result, they are attracted to positive objects and repelled from negative objects, including other electrons (illustrated below). To minimize repulsion, each electron is capable of staking out a "territory" and "defending" itself from other electrons.

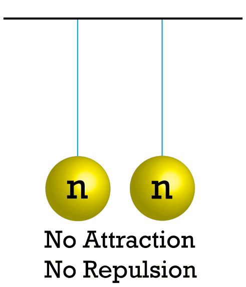


Protons are another type of subatomic particle found in atoms. They have a positive charge, so they are attracted to negative objects and repelled from positive objects. Again, this means that protons repel each other (illustrated below). However, unlike electrons, protons are forced to group together into one big clump, even though they repel each other. Protons are bound together by what are termed **strong nuclear forces**. These forces are responsible for binding the atomic nuclei together, allowing the protons to form a dense, positively charged center.



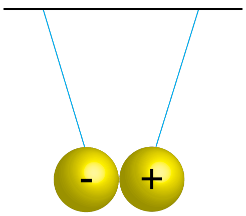
There is a third subatomic particle known as a **neutron**. Rutherford proposed the existence of a neutral particle along with the approximate mass of a proton, but it wasn't until years later that someone proved the existence of the neutron. A physicist named James Chadwick observed that when beryllium was bombarded with alpha particles, it emitted an unknown radiation that had approximately the same mass as a proton, but the radiation had no electrical charge. Chadwick was able to prove that these beryllium emissions contained a neutral particle – Rutherford's neutron.

As you might have already guessed from its name, the neutron is neutral. In other words, it has no charge and is therefore neither attracted to nor repelled from other objects. Neutrons are in every atom (with one exception), and they're bound together with other neutrons and protons in the atomic nucleus. Again, the binding forces that help to keep neutrons fastened into the nucleus are known as strong nuclear forces.



Since neutrons are neither attracted to nor repelled from objects, they don't really interact with protons or electrons beyond being bound into the nucleus with the protons. Protons and electrons, however, do interact. Using what you know about protons and electrons, what do you think will happen when an

electron approaches a proton? Will the two subatomic particles be attracted to each other or repelled from each other? Here's a hint: "opposites attract, likes repel." Since electrons and protons have opposite charges (one negative, the other positive), you'd expect them to be attracted to each other, as illustrated below.



Even though electrons, protons, and neutrons are all types of subatomic particles, they are not all the same size. When comparing the masses of electrons, protons, and neutrons, you will find that electrons have an extremely small mass compared to the masses of either protons or neutrons (see **Figure** 1.11). On the other hand, the masses of protons and neutrons are fairly similar, with the mass of a neutron being slightly greater than the mass of a proton. Because protons and neutrons are so much more massive than electrons, almost all of the atomic mass in any atom comes from the nucleus, which is where all of the neutrons and protons are located.

Figure 1.11: Electrons are much smaller than protons or neutrons. How much smaller? If an electron was the size of a penny, a proton or a neutron would have the mass of a large bowling ball!

Table 1.1 gives the properties and locations of electrons, protons, and neutrons. The third column shows the masses of the three subatomic particles in grams. The second column, however, shows the masses of the three subatomic particles in amu, or atomic mass units. An **atomic mass unit (amu)** is defined as one-twelfth the mass of a carbon-12 atom (a carbon that has 6 protons and 6 neutrons). Atomic mass units are useful because, as you can see, the mass of a proton and the mass of a neutron are almost exactly 1.0 in this unit system. The **dalton** is equivalent to the atomic mass unit, with the two terms being different names for the same measure. The two terms are often used interchangeably, and both will be used throughout this text.

Table 1.1: **Subatomic Particles, Properties, and Location**

Particle	Relative Mass (amu)	Mass in Grams (g)	Electric Charge	Location
electron	$\frac{1}{1840}$	9.109383×10^{-28}	-1	outside nucleus
proton	1	$1.6726217 \times 10^{-24}$	$+1$	nucleus
neutron	1	$1.6749273 \times 10^{-24}$	0	nucleus

In addition to mass, another important property of subatomic particles is the charge. The fourth column in **Table** 1.1 shows the charges of the three subatomic particles. You already know that neutrons are neutral and thus have no charge at all. Therefore, we say that neutrons have a charge of zero. What about electrons and protons? Electrons are negatively charged and protons are positively charged, but the positive charge on a proton is exactly equal in magnitude (magnitude means "absolute value") to the negative charge on an electron. You may recall that Millikan discovered that the magnitude of the charge on a single electron is 1.6×10^{-19} C (coulomb), which means that the magnitude of the charge on a proton is also 1.6×10^{-19} C. In other words, a neutral atom must have exactly one electron for every proton. If a neutral atom has 1 proton, it must have 1 electron. If a neutral atom has 2 protons, it must have 2 electrons. If a neutral atom has 10 protons, it must have 10 electrons. Do you get the idea?

For a short animation demonstrating the properties of the electron using a cathode ray tube **(1h)**, see http://www.youtube.com/watch?v=4QAzu6fe8rE video (3:46).

Figure 1.12: (Watch Youtube Video)
http://www.ck12.org/flexbook/embed/view/360

Atomic Number and Mass Number

Scientists can distinguish between different elements by counting the number of protons. If an atom has only one proton, we know it's an atom of the element hydrogen. An atom with two protons is always an atom of the element helium. When scientists count four protons in an atom, they know it's a beryllium atom. An atom with three protons is a lithium atom, an atom with five protons is a boron atom, an atom with six protons is a carbon atom... the list goes on (see **Figure** 1.13 for more examples).

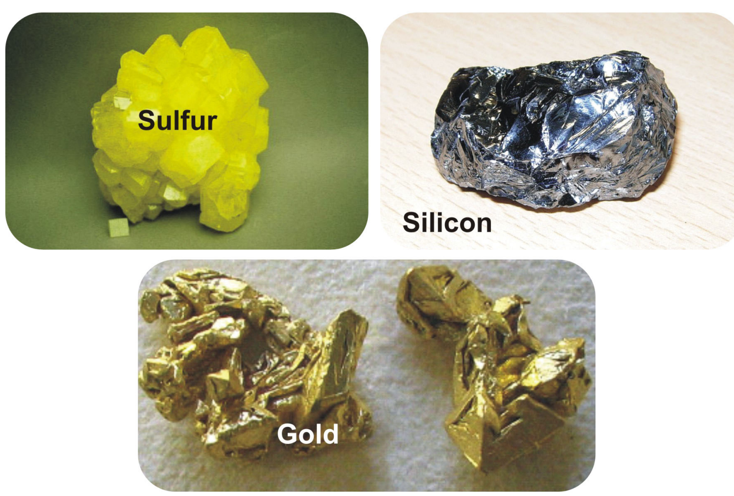


Figure 1.13: How would you distinguish these three elements? You can

Since an atom of one element can be distinguished from an atom of another element by the number of protons in the nucleus, scientists are always interested in this number and how this number differs between different elements. Therefore, scientists give this number a special name and a special symbol. An element's **atomic number (*Z*)** is equal to the number of protons in the nuclei of any of its atoms. The periodic table gives the atomic number of each element. The atomic number is a whole number usually written above the chemical symbol of each element in the table. The atomic number for hydrogen is $Z = 1$ because every hydrogen atom has 1 proton. The atomic number for helium is $Z = 2$ because every helium atom has 2 protons. What is the atomic number of carbon? (*Answer*: Carbon has 6 protons, so the atomic number for carbon is $Z = 6$.)

Since neutral atoms have to have one electron for every proton, an element's atomic number also tells you how many electrons are in a neutral atom of that element. For example, hydrogen has atomic number $Z = 1$. This means that an atom of hydrogen has one proton and, if it's neutral, one electron. Gold, on the other hand, has atomic number $Z = 79$, which means that a neutral atom of gold has 79 protons and 79 electrons.

The **mass number (*A*)** of an atom is the total number of protons and neutrons in its nucleus. Why do you think that the mass number includes protons and neutrons, but not electrons? You know that most of the mass of an atom is concentrated in its nucleus and that the mass of an electron is very, very small compared to the mass of either a proton or a neutron (like the mass of a penny compared to the mass of a bowling ball). By counting the number of protons and neutrons, scientists will have a very close approximation of the total mass of an atom.

mass number A = (number of protons) + (number of neutrons)

An atom's mass number is very easy to calculate once you know the number of protons and neutrons in the atom. Notice that the mass number is not the same as the mass of the atom. You can easily relate the mass number to the mass by recalling that both protons and neutrons have a relative mass of 1 amu.

Example:

What is the mass number of an atom that contains 3 protons and 4 neutrons?

mass number $A = (3) + (4) = 7$
mass number $A =$ (number of protons) + (number of neutrons)
(number of neutrons) = 4
(number of protons) = 3

Example:

What is the mass number of an atom of helium that contains 2 neutrons?

mass number $A = (2) + (2) = 4$
mass number $A =$ (number of protons) + (number of neutrons)
(number of neutrons) = 2
(number of protons) = 2 (Remember that an atom of helium always has 2 protons.)

This video summarizes the concept of the atom and to the organization of the periodic table **(1a, 1e)**: http://www.youtube.com/watch?v=1xSQlwWGT8M (21:05).

Isotopes and Atomic Mass

Unlike the number of protons, which is always the same for all atoms of the same element, the number of neutrons can be different. Atoms of the same element with different numbers of neutrons are known as

Figure 1.14: (Watch Youtube Video)
http://www.ck12.org/flexbook/embed/view/361

isotopes. Since the isotopes of any given element all contain the same number of protons, they have the same atomic number. However, since the isotopes of a given element contain different numbers of neutrons, different isotopes have different mass numbers. The following two examples should help to clarify this point.

Example:

What is the atomic number (Z) and the mass number (A) of an isotope of lithium containing 3 neutrons? A lithium atom contains 3 protons in its nucleus.

mass number $A = (3) + (3) = 6$
mass number $A =$ (number of protons) + (number of neutrons)
number of neutrons $= 3$
atomic number $Z =$ number of protons $= 3$

Example:

What is the atomic number (Z) and the mass number (A) of an isotope of lithium containing 4 neutrons? A lithium atom contains 3 protons in its nucleus.

mass number $A = (3) + (4) = 7$
mass number $A =$ (number of protons) + (number of neutrons)
number of neutrons $= 4$
atomic number $Z =$ number of protons $= 3$

Notice that because the lithium atom always has 3 protons, the atomic number for lithium is always $Z = 3$. The mass number, however, is $A = 6$ for the isotope with 3 neutrons, and $A = 7$ for the isotope with 4 neutrons. In nature, only certain isotopes exist. For instance, lithium exists as an isotope with 3 neutrons and as an isotope with 4 neutrons, but it doesn't exists as an isotope with 2 neutrons or as an isotope with 5 neutrons.

This whole discussion of isotopes brings us back to Dalton's atomic theory. According to Dalton, atoms of a given element are identical. But if atoms of a given element can have different numbers of neutrons, then they can have different masses as well. How did Dalton miss this? It turns out that elements found in nature exist as uniform mixtures with a constant ratio of their naturally occurring isotopes. In other words, a piece of lithium always contains both types of naturally occurring lithium (the type with 3 neutrons and the type with 4 neutrons). Moreover, it always contains the two in the same relative amounts (or "relative abundances"). In a chunk of lithium, 93% will always be lithium with 4 neutrons, while the remaining 7% will always be lithium with 3 neutrons.

Unfortunately, Dalton always experimented with large chunks of an element – chunks that contained all of the naturally occurring isotopes of that element. As a result, when he performed his measurements, he

was actually observing the averaged properties of all the different isotopes in the sample. Luckily, aside from having different masses, most other properties of different isotopes are similar.

Knowing about the different isotopes is important when it comes to calculating atomic mass. The **atomic mass** (sometimes referred to as atomic weight) of an element is the weighted average mass of the atoms in a naturally occurring sample of the element. Atomic mass is typically reported in atomic mass units. You can calculate the atomic mass of an element, provided you know the relative abundances the element's naturally occurring isotopes and the masses of those different isotopes. The examples below show how this calculation is done.

Example:

Boron has two naturally occurring isotopes. In a sample of boron, 20% of the atoms are B-10, which is an isotope of boron with 5 neutrons and a mass of 10 amu. The other 80% of the atoms are B-11, which is an isotope of boron with 6 neutrons and a mass of 11 amu. What is the atomic mass of boron?

Solution:

To do this problem, we will calculate 20% of the mass of B-10, which is how much the B-10 isotope contributes to the "average boron atom." We will also calculate 80% of the mass of B-11, which is how much the B-11 isotope contributes to the "average boron atom."

Step One: Convert the percentages given in the question into their decimal forms by dividing each percentage by 100%:

Decimal form of $80\% = 0.80$
Decimal form of $20\% = 0.20$

Step Two: Multiply the mass of each isotope by its relative abundance (percentage) in decimal form:

80% of the mass of B-11 $= 0.80 \times 11$ amu $= 8.8$ amu
20% of the mass of B-10 $= 0.20 \times 10$ amu $= 2.0$ amu

Step Three: Find the total mass of the "average atom" by adding together the contributions from the different isotopes:

Total mass of average atom $= 2.0$ amu $+ 8.8$ amu $= 10.8$ amu

The mass of an average boron atom, and thus boron's atomic mass, is 10.8 amu.

Example:

Neon has three naturally occurring isotopes. In a sample of neon, 90.48% of the atoms are Ne-20, which is an isotope of neon with 10 neutrons and a mass of 19.99 amu. Another 0.27% of the atoms are Ne-21, which is an isotope of neon with 11 neutrons and a mass of 20.99 amu. The final 9.25% of the atoms are Ne-22, which is an isotope of neon with 12 neutrons and a mass of 21.99 amu. What is the atomic mass of neon?

Solution:

To do this problem, we will calculate 90.48% of the mass of Ne-20, which is how much Ne-20 contributes to the "average neon atom." We will also calculate 0.27% of the mass of Ne-21 and 9.25% of the mass of Ne-22, which are how much the Ne-21 and the Ne-22 isotopes contribute to the "average neon atom" respectively.

Step One: Convert the percentages given in the question into their decimal forms by dividing each percentage by 100%:

Decimal form of 9.25% = 0.0925
Decimal form of 0.27% = 0.0027
Decimal form of 90.48% = 0.9048

Step Two: Multiply the mass of each isotope by its relative abundance (percentage) in decimal form:

9.25% of the mass of Ne-22 = 0.0885×22.00 amu = 2.04 amu
0.27% of the mass of Ne-21 = 0.0027×21.00 amu = 0.057 amu
90.48% of the mass of Ne-20 = 0.9048×20.00 amu = 18.10 amu

Step Three: Find the total mass of the "average atom" by adding together the contributions from the different isotopes:

Total mass of average atom = 18.10 amu + 0.057 amu + 2.04 amu = 20.20 amu

The mass of an average neon atom, and thus neon's atomic mass, is 20.20 amu.

The periodic table gives the atomic mass of each element. The atomic mass is a number that usually appears below the element's symbol in each square. Notice that atomic mass of boron (symbol B) is 10.8 and the atomic mass of neon (symbol Ne) is 20.18, both which are very close to what we calculated in our examples. Take time to notice that not all periodic tables have the atomic number above the element's symbol and the atomic mass below it. If you are ever confused, remember that the atomic number should always be the smaller of the two and will be a whole number, while the atomic mass should always be the larger of the two. (The atomic mass must include both the number of protons and the average number of neutrons.)

Lesson Summary

- Electrons are a type of subatomic particle with a negative charge, so electrons repel each other but are attracted to protons.
- Protons are a type of subatomic particle with a positive charge, so protons repel each other but are attracted to electrons. Protons are bound together in an atom's nucleus as a result of strong nuclear forces.
- Neutrons are a type of subatomic particle with no charge (they're neutral). Like protons, neutrons are bound into the atom's nucleus as a result of strong nuclear forces.
- Protons and neutrons have approximately the same mass and are both much more massive than electrons (approximately 2,000 times as massive as an electron).
- The positive charge on a proton is equal in magnitude to the negative charge on an electron. As a result, a neutral atom must have an equal number of protons and electrons.
- Each element has a unique number of protons. An element's atomic number (Z) is equal to the number of protons in the nuclei of any of its atoms.
- The mass number (A) of an atom is the sum of the protons and neutrons in the atom
- Isotopes are atoms of the same element (same number of protons) that have different numbers of neutrons in their atomic nuclei.
- An element's atomic mass is the average mass of one atom of that element. An element's atomic mass can be calculated provided the relative abundances of the element's naturally occurring isotopes and the masses of those isotopes are known.
- The periodic table is a convenient way to summarize information about the different elements. In addition to the element's symbol, most periodic tables will also contain the element's atomic number and the element's atomic mass.

Further Reading / Supplemental Links

This website has a video called "Atomic Structure: The Nucleus" available.

- http://videos.howstuffworks.com/hsw/5806-atomic-structure-the-nucleus-video.htm

Review Questions

1. Decide whether each of the following statements is true or false.
   (a) The nucleus of an atom contains all of the protons in the atom.
   (b) The nucleus of an atom contains all of the neutrons in the atom.
   (c) The nucleus of an atom contains all of the electrons in the atom.
   (d) Neutral atoms of a given element must contain the same number of neutrons.
   (e) Neutral atoms of a given element must contain the same number of electrons.
2. Match the subatomic property with its description in **Table** 1.2.

Table 1.2: **Table for Problem 2**

Subatomic Particle	**Characteristics**
i. electron	a. has an atomic charge of $+1\ e$
ii. neutron	b. has a mass of 9.109383×10^{-28} grams
iii. proton	c. is neither attracted to nor repelled from charged objects

3. Arrange the electron, proton, and neutron in order of decreasing mass.
4. Indicate which of the following statements is true or false.
   (a) An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
   (b) The symbol for an element's atomic number is A.
   (c) A neutral atom with $Z = 4$ must have 4 electrons.
   (d) A neutral atom with $A = 4$ must have 4 electrons.
   (e) An atom with 7 protons and 7 neutrons will have $A = 14$.
   (f) An atom with 7 protons and 7 neutrons will have $Z = 14$.
   (g) A neutral atom with 7 electrons and 7 neutrons will have $A = 14$.
5. Use the periodic table to find the symbol for the element with:
   (a) 44 electrons in a neutral atom.
   (b) 30 protons.
   (c) $Z = 36$.
   (d) an atomic mass of 14.007 amu.
6. When will the mass number (A) of an atom be:
   (a) bigger than the atomic number (Z) of the atom?
   (b) smaller than the atomic number (Z) of the atom?
   (c) equal to the atomic number (Z) of the atom?
7. In **Table** 1.3, Column 1 contains data for five different elements. Column 2 contains data for the same five elements but with different isotopes of those elements. Match the columns by connecting isotopes of the same element.

Table 1.3: **Table for Problem 7**

Column 1	**Column 2**
a. an atom with 2 protons and 1 neutron	i. a C (carbon) atom with 6 neutrons
b. a Be (beryllium) atom with 5 neutrons	ii. an atom with 2 protons and 2 neutrons
c. an atom with $Z = 6$ and $A = 13$	iii. an atom with $Z = 7$ and $A = 15$
d. an atom with 1 proton and $A = 1$	iv. an atom with $A = 2$ and 1 neutron
e. an atom with $Z = 7$ and 7 neutrons	v. an atom with $Z = 4$ and 6 neutrons

8. Match the following isotopes with their respective mass numbers in **Table** 1.4.

Table 1.4: **Table for Problem 8**

Column 1	**Column 2**
(a) an atom with $Z = 17$ and 18 neutrons	i. 35
(b) an H atom with no neutrons	ii. 4
(c) A He atom with 2 neutrons	iii. 1
(d) an atom with $Z = 11$ and 11 neutrons	iv. 23
(e) an atom with 11 neutrons and 12 protons	v. 22

9. Match the following isotopes with their respective atomic numbers in **Table** 1.5.

Table 1.5: **Table for Problem 9**

Column 1	**Column 2**
(a) a B (boron) atom with $A = 10$	i. 8
(b) an atom with $A = 10$ and 6 neutrons	ii. 2
(c) an atom with 3 protons and 3 neutrons	iii. 3
(d) an oxygen atom	iv. 4
(e) an atom with $A = 4$ and 2 neutrons	v. 5

10. Answer the following questions:
    (a) What's the mass number of an atom that contains 13 protons and 13 neutrons?
    (b) What's the mass number of an atom that contains 24 protons and 30 neutrons?
11. Answer the following questions:
    (a) What's the mass number of the isotope of manganese (Mn) containing 28 neutrons?
    (b) What's the mass number of the isotope of calcium (Ca) containing 20 neutrons?
12. Answer the following questions:
    (a) What's the atomic number of an atom that has 30 neutrons, and a mass number of $A = 70$?
    (b) What's the atomic number of an atom with 14 neutrons, if the mass number of the atom is $A = 28$?
13. Answer the following questions:
    (a) What's the mass number of a neutral atom that contains 7 protons and 7 neutrons?

(b) What's the mass number of a neutral atom that contains 7 electrons and 7 neutrons?
(c) What's the mass number of a neutral atom that contains 5 protons, 5 electrons, and 6 neutrons?
(d) What's the mass number of a neutral atom that contains 3 electrons and 4 neutrons?

14. Answer the following questions:
    (a) What element has 32 neutrons in an atom with mass number $A = 58$?
    (b) What element has 10 neutrons in an atom with mass number $A = 19$?
15. Copper has two naturally occurring isotopes. 69.15% of copper atoms are Cu-63 and have a mass of 62.93 amu. The other 30.85% of copper atoms are Cu-65 and have a mass of 64.93 amu. What is the atomic mass of copper?