Chapter 2: The Structure of the Atom

2.1: Early Ideas of Atoms

Objectives

- Give a short history of the concept of the atom.
- Describe the contributions of Democritus and Dalton to atomic theory.
- Summarize Dalton's atomic theory and explain its historical development.

Introduction

You learned earlier how all matter in the universe is made out 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. It has always been difficult to convince people of the existence of things that are too small to see. We will spend some time considering the evidence (observations) that convince scientists of the existence of atoms.

Democritus and the Greek Philosophers

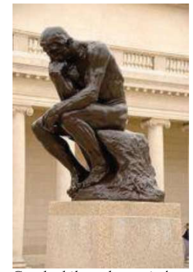
Before we discuss the experiments and evidence that have, over the years, convinced scientists that matter is made up of atoms, it's only fair to give credit to the man who proposed "atoms" in the first place. About 2,500 years ago, early Greek philosophers believed the entire universe was a single, 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."

Democritus was known as "The Laughing Philosopher." It's a good thing he liked to laugh, because most other philosophers were laughing at his theories.

One of the first people to propose "atoms" was a man known as Democritus. As an alternative to the beliefs of the Greek philosophers, he suggested that **atomos**, or atomon – tiny, indivisible, solid objects - make 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, suggesting 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 different types of atomos. According to Democritus, other characteristics, like color and taste, did not reflect properties of the atomos themselves, but rather, resulted from the different ways in which the atomos were combined and connected to one another.

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.



Greek philosophers tried to understand the nature of the world through reason and logic but not through experiment and observation.

So how could the Greek philosophers have known that Democritus had a good idea with his theory of "atomos?" It would have taken some careful observation and a few simple experiments. Now you might wonder why Greek philosophers didn't perform any experiments to actually test Democritus' theory. The problem, of course, was that Greek philosophers didn't believe in experiments at all. Remember, Greek philosophers didn't trust their senses, they only trusted the reasoning power of the mind.

The early Greek philosophers tried to understand the nature of the world through reason and logic, but not through experiment and observation. As a result, they had some very interesting ideas, but they felt no need to justify their ideas based on life experiences. In a lot of ways, you can think of the Greek philosophers as being "all thought and no action." It's truly amazing how much they achieved using their minds, but because they never performed any experiments, they missed or rejected a lot of discoveries that they could have made otherwise. Greek philosophers dismissed Democritus' theory entirely. Sadly, it took over two millennia before the theory of atomos (or "atoms," as they're known today) was fully appreciated.

Dalton's Atomic Theory

Although the concept of atoms is now widely accepted, this wasn't always the case. Scientists didn't always believe that everything was composed of small particles called atoms. The work of several scientists and their experimental data gave evidence for what is now called the atomic theory.

In the late 1700's, Antoine Lavoisier, a French scientist, experimented with the reactions of many metals. He carefully measured the mass of a substance before reacting and again measured the mass after a reaction had occurred in a closed system (meaning that nothing could enter or leave the container). He found that no matter what reaction he looked at, the mass of the starting materials was always equal to the mass of the ending materials. This is now called the **law of conservation of mass**. This went contrary to what many scientists at the time thought. For example, when a piece of iron rusts, it appears to gain mass. When a log is burned, it appears to lose mass. In these examples, though, the reaction does not take place in a closed container and substances, such as the gases in the air, are able to enter or leave. When iron rusts, it is combining with oxygen in the air, which is why it seems to gain mass. What Lavoisier found was that no mass was actually being gained or lost. It was coming from the air. This was a very important first step in giving evidence for the idea that everything is made of atoms. The atoms (and mass) are not being created or destroyed. The atoms are simply reacting with other atoms that are already present.

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, like water, 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 special. In fact, it turned out that the reactants in every chemical reaction reacted in the same proportions by mass. This observation is summarized in the **law of definite proportions.** Take, for example, nitrogen and hydrogen, which react to produce ammonia. 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.

At the same time that scientists were finding this pattern out, a man named John Dalton was experimenting with several reactions in which the reactant elements formed more than one type of product, 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 can make A with 3 grams of carbon and 4 grams of oxygen, B can be made with the same 3 grams of carbon, but with 8 grams oxygen. Dalton asked himself – why does B require 2 times as much oxygen as A? Why not 1.21 times as much oxygen, or 0.95 times as much oxygen? Why a whole number like 2?

Unlike the Greek philosophers, John Dalton believed in both logical thinking and experimentation.

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. Similarly, E always required exactly four times as much oxygen as C. Once again, Dalton noticed that small whole numbers (2 and 4) seemed to be the rule. This observation came to be known as the **law of multiple proportions**.

Dalton thought about his results and tried to find some theory that would explain it, as well as a theory that would explain the Law of Conservation of Mass (mass is neither created nor destroyed, or the mass you have at the beginning is equal to the mass at the end of a change). 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' original idea). Dalton used this idea to generate what is now known as **Dalton's Atomic Theory** which stated the following:

1. Matter is made of tiny particles called atoms.
2. Atoms are indivisible (can't be broken into smaller particles). 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" – new, complex particles. In a given compound, however, the different types of atoms are always present in the same relative numbers.

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' "atomos" theory because they felt it was illogical.
- Dalton used observations about the ratios in which elements will react to combine 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" – new complex particles. In a given compound, however, the different types of atoms are always present in the same relative numbers.

Vocabulary

- Atom: Democritus' word for the tiny, indivisible, solid objects that he believed made up all matter in the universe
- Dalton's Atomic Theory: the first scientific theory to relate chemical changes to the structure, properties, and behavior of the atom

Further Reading / Supplemental Links

- To see a video documenting the early history of the concept of the atom, go to http://www.uen.org/dms/. Go to the k-12 library. Search for "history of the atom". Watch part 01. (you can get the username and password from your teacher)
- Vision Learning: From Democritus to Dalton: http://visionlearning.com/library/module_viewer.php?c3=&mid=49&l=

2.1: Review Questions

1) *(Multiple choice)* 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.

2) Democritus and Dalton both suggested that all matter was composed of small particles, called atoms. What is the greatest advantage Dalton's Atomic Theory had over Democritus'?

3) 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?

2.2: Further Understanding of the Atom

Objectives

- Explain the observations that led to Thomson's discovery of the electron.
- Describe Thomson's "plum pudding" mode of the atom and the evidence for it
- 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 altered the "plum pudding" model.
- Draw a diagram of the Rutherford model of the atom and label the nucleus and the electron cloud.

Introduction

Dalton's Atomic Theory held up well to a lot of the different chemical experiments that scientists performed to test it. In fact, for almost 100 years, it seemed as if Dalton's Atomic Theory was the whole truth. However, in 1897, a scientist named J. J. Thomson conducted some research that suggested that Dalton's Atomic Theory wasn't the entire story. As it turns out, Dalton had 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 during a chemical reaction, atoms are merely rearranged; 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. We'll talk about the discoveries of these subatomic particles next.

J.J. Thomson conducted experiments that suggested that Dalton's atomic theory wasn't telling the entire story.

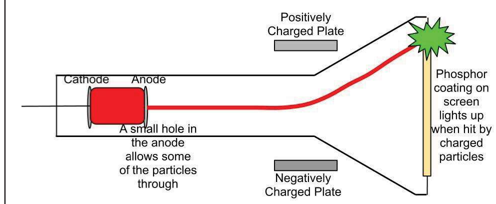
Thomson's Plum Pudding Model

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.

The figure shows a basic diagram of a cathode ray tube like the one J. J. 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 by 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. J. J. 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 J. J. Thomson 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, J. J. 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, J. J. Thomson 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.

J. J. 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 of the cathode ray tube, as shown in Figure 3. 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. Using the "opposite charges attract, like charges repel" rule, J. J. Thomson argued that if the cathode rays were attracted to the positively charged metal plate and repelled from the negatively charged metal plate, they themselves must have a negative charge!



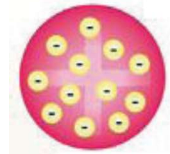
Thomson's experiment with cathode rays found that the ray moved away from negatively charged plates and toward positively charges plates. What does this say about the charge of the ray?
CC – Tracy Poulsen

J. J. Thomson then did some rather complex experiments with magnets, and used his results to prove that cathode rays were not only negatively charged, but 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. But there was a problem. According to J. J. Thomson's measurements, either these cathode rays had a ridiculously high charge, or else had 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? J. J. Thomson made a radical proposal: maybe atoms are divisible. J. J. Thomson 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, J. J. Thomson is credited with the discovery of the electron.

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 whole lot of electrons. But there's another, even bigger problem: electrons are negatively charged. Therefore, if atoms were made entirely out of electrons, atoms would be negatively charged themselves… and that would mean all matter was negatively charged as well. Of course, matter isn't negatively charged. In fact, most matter is what we call neutral – it has no charge at all. If matter is composed of atoms, and atoms are composed of negative electrons, how can matter be neutral? The only possible explanation is that atoms consist of more than just electrons. Atoms must also contain some type of positively charged material that balances the negative charge on the electrons. Negative and positive charges of equal size cancel each other out, just like negative and positive numbers of equal size. What do you get if you add +1 and -1? You get 0, or nothing. That's true of numbers, and that's also true of charges. If an atom contains an electron with a -1 charge, but also some form of material with a +1 charge, overall the atom must have a (+1) + (-1) = 0 charge – in other words, the atom must be neutral, or have no charge at all.

Based on the fact that atoms are neutral, and based on J. J. 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 realized that atoms had to contain positive material as well – otherwise they wouldn't be neutral overall. 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. 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. As a result, Thomson theorized that the positive material in the atom must form something like the "batter" in a plum pudding, while the negative electrons must be scattered through this "batter." (If you've never seen or tasted a plum pudding, you can think of a chocolate chip cookie instead. 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.)



Thomson's plum pudding model was much like a chocolate chip cookie. Notice how the chocolate chips are the negatively charged electrons, while the positive charge is spread throughout the entire

Notice 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'd expect it would be easy to pick out the electrons, but a lot harder to pick out the positive material.

J.J. Thomson had measured the charge to mass ratio of the electron, but had been unable to accurately measure the charge on the electron. With his oil drop experiment, Robert Millikan was able to accurately measure the charge of the electron. When combined with the charge to mass ratio, he was able to calculate the mass of the electron. What Millikan did was to put a charge on tiny droplets of oil and measured 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, the charge on a single electron.

Rutherford's Nuclear Model

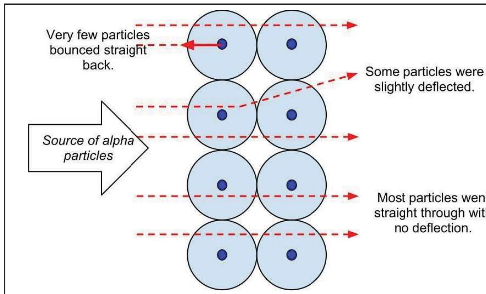
Everything about Thomson's experiments suggested the "plum pudding" model was correct – but according to the scientific method, any new theory or model should be tested by further experimentation and observation. In the case of the "plum pudding" model, it would take a man named Ernest Rutherford to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

Ernest Rutherford

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, they were, however, used to conduct some very interesting experiments.

Ernest Rutherford was fascinated by all aspects of alpha particles. For the most part, though, he seemed to view alpha particles as tiny bullets that he could use to fire at all kinds of different materials. One experiment in particular, however, surprised Rutherford, and everyone else.

Rutherford found that when he fired alpha particles at a very thin piece of gold foil, an interesting thing happened. Almost all of the alpha particles went straight through the foil as if they'd hit nothing at all. This was what he expected to happen. If Thomson's model was accurate, there was nothing hard enough for these small particles to hit that would cause any change in their motion.

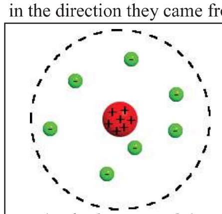


Ernest Rutherford's Gold Foil Experiment in which alpha particles were shot at a piece of gold foil. Most of the particles went straight through, but some bounced straight back, indicating they were hitting a very small, very dense particle in the atom.
CC – Tracy Poulsen

Every so often, though, one of the alpha particles would be deflected slightly as if it had bounced off of 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.

Rutherford thought that these experimental results were rather odd. Rutherford described firing alpha particles at gold foil like shooting a high-powered rifle at tissue paper. Would you ever expect the bullets to hit the tissue paper and bounce back at you? Of course not! The bullets would break through the tissue paper and keep on going, almost as if they'd hit nothing at all. That's what Rutherford had expected would happen when he fired alpha particles at the gold foil. Therefore, the fact that most alpha particles passed through didn't shock him. On the other hand, how could he explain the alpha particles that got deflected? Furthermore, how could he explain the alpha particles that bounced right back as if they'd 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, in fact, distributed like the batter in plum pudding, but rather, 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 alpha particles passed through the gold foil without hitting anything at all. Once in a while, though, the alpha particles would actually collide with a gold nucleus, causing the alpha particles to be deflected, or even to bounce right back in the direction they came from.



Rutherford suggested that electrons surround a central nucleus.
(Obtained from: http://upload.wikimedia.org/wikipedia/commons/7/7d/Rutherfordsches_Atommodell.png)

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 wasn't distributed like batter, but rather, 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? And 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, but at the time, there wasn't much evidence to suggest exactly how the electrons were held in the atom.

Despite the problems and questions associated with Rutherford's experiments, his work with alpha particles definitely seemed to point to the existence of an atomic "nucleus." Between J. J. Thomson, who discovered the electron, and Rutherford, who suggested that the positive charges in an atom were concentrated at the atom's center, the 1890s and early 1900s saw huge steps in understanding the atom at the "subatomic" (or 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, it was clear that an atom contains negatively charged electrons and a nucleus containing positive

charges. In the next section, we'll look more carefully at the structure of the nucleus, and 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. It turns out 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.

Vocabulary

- Subatomic particles: particles that are smaller than the atom
- Electron: a negatively charged subatomic particle
- Proton: a positively charged subatomic particle
- Nucleus: the small, dense center of the atom

Further Reading / Supplemental Material

- A short history of the changes in our model of the atom, an image of the plum pudding model, and an animation of Rutherford's experiment can be viewed at Plum Pudding and Rutherford Page (http://www.newcastle-schools.org.uk/nsn/chemistry/Radioactivity/Plub%20Pudding%20and%20Rutherford%20Page.htm).
- To see a video documenting the early history of the concept of the atom, go to http://www.uen.org/dms/. Go to the k-12 library. Search for "history of the atom". Watch part 02. (you can get the username and password from your teacher)
- Vision Learning: The Early Days (Thomson, etc) http://visionlearning.com/library/module_viewer.php?mid=50&l=&c3=
- Discovery of Electron (YouTube): http://www.youtube.com/watch%3Fv%3DIdTxGJjA4Jw
- Thomson's Experiment: http://www.aip.org/history/electron/jjthomson.htm
- Discovery of Atomic Nucleus (YouTube): http://www.youtube.com/watch%3Fv%3DwzALbzTdnc8
- Rutherford's Experiment: http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/ruther14.swf

2.2: Review Questions

Decide whether each of the following statements is true or false.

1) Electrons (cathode rays) are positively charged.

2) Electrons (cathode rays) can be repelled by a negatively charged metal plate.
3) J.J. Thomson is credited with the discovery of the electron.
4) The plum pudding model is the currently accepted model of the atom

#5-11: Match each conclusion regarding subatomic particles and atoms with the observation/data that supports it.

Conclusion	Observations
5) All atoms have electrons	a. Most alpha particles shot at gold foil go straight through, without any change in their direction.
6) Atoms are mostly empty space.	b. A few alpha particles shot at gold foil bounce in the opposite direction.
7) Electrons have a negative charge	c. Some alpha particles (with positive charges) when shot through gold foil bend away from the gold.
8) The nucleus is positively charged	d. No matter which element Thomson put in a cathode ray tube, the same negative particles with the same properties (such as charge & mass) were ejected.
9) Atoms have a small, dense nucleus	e. The particles ejected in Thomson's experiment bent away from negatively charged plates, but toward positively charged plates.

10) What is the name given to the tiny clump of positive material at the center of an atom?

11) Electrons are ______ negatively charged metals plates and ______ positively charged metal plates.

Consider the following two paragraphs for #12-14

Scientist 1: Although atoms were once regarded as the smallest part of nature, they are composed of even smaller particles. All atoms contain negatively charged particles, called electrons. However, the total charge of any atom is zero. Therefore, this means that there must also be positive charge in the atom. The electrons sit in a bed of positively charged mass.

Scientist 2: It is true that atoms contain smaller particles. However, the electrons are not floating in a bed of positive charge. The positive charge is located in the central part of the atom, in a very small, dense mass, called a nucleus. The electrons are found outside of the nucleus.

12) What is the main dispute between the two scientists' theories?
13) Another scientist was able to calculate the exact charge of an electron to be -1.6×10^{-19} C. What effect does this have on the claims of Scientist 1? (Pick one answer)
    a) Goes against his claim
    b) Supports his claim
    c) Has no effect on his claim.
14) If a positively charged particle was shot at a thin sheet of gold foil, what would the second scientist predict to happen?

2.3: Protons, Neutrons, and Electrons in Atoms

Objectives

- Describe the locations, charges, and masses and the three main subatomic particles.
- Define atomic number.
- Describe the size of the nucleus in relation to the size of the atom.
- 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.

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. We have been talking about the electron in great detail, but there are two other particles of interest to use: protons and neutrons. In this section, we'll look at the atom a little more closely.

Protons, Electrons, and Neutrons

We already learned that J.J. Thomson discovered a negatively charged particle, called the **electron**. Rutherford proposed that these electrons orbit a positive nucleus. In subsequent experiments, he found that there is a smaller positively charged particle in the nucleus which is called a **proton**. There is a third subatomic particle, known as a neutron. Ernest Rutherford proposed the existence of a neutral particle, with the approximate mass of a proton. Years later, James Chadwick proved that the nucleus of the atom contains this neutral particle that had been proposed by Ernest Rutherford. Chadwick observed that when beryllium is bombarded with alpha particles, it emits an unknown radiation that has approximately the same mass as a proton, but no electrical charge. Chadwick was able to prove that the beryllium emissions contained a neutral particle - Rutherford's neutron.

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

As you might have already guessed from its name, the neutron is neutral. In other words, it has no charge whatsoever, 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.

Before we move on, we must discuss how the different types of subatomic particles interact with each other. When it comes to neutrons, the answer is obvious. 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).

Even though electrons, protons, and neutrons are all types of subatomic particles, they are not all the same size. When you compare the masses of electrons, protons and neutrons, what you find is that electrons have an extremely small mass, compared to either protons or neutrons. On the other hand, the masses of protons and neutrons are fairly similar, although technically, the mass of a neutron is slightly larger than the mass of a proton. Because protons and neutrons are so much more massive than electrons, almost all of the mass of any atom comes from the nucleus, which contains all of the neutrons and protons.

The table shown 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 "atomic mass units". An **atomic mass unit (amu)** is defined as one-twelfth the mass of a carbon-12 atom. Atomic mass units (amu) 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.

Sub-Atomic Particles, Properties and Location

Particle	Relative Mass (amu)	Electric Charge	Location
electron	$\frac{1}{1840}$	-1	outside the nucleus
proton	1	+1	nucleus
neutron	1	0	nucleus

In addition to mass, another important property of subatomic particles is their charge. 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? You know that electrons are negatively charged and protons are positively charged, but what's amazing is that the positive charge on a proton is exactly equal in magnitude (magnitude means "absolute value" or "size when you ignore positive and negative signs") to the negative charge on an electron. The third column in the table shows the charges of the three subatomic particles. Notice that the charge on the proton and the charge on the electron have the same magnitude.

Negative and positive charges of equal magnitude cancel each other out. This means that the negative charge on an electron perfectly balances the positive charge on the proton. 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. You get the idea. In order to be neutral, an atom must have the same number of electrons and protons.

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 a hydrogen atom. An atom with two protons is always a helium atom. If 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.

It is difficult to find qualities that are different from each element and distinguish on element from another. Each element, however, does have a unique number of protons. Sulfur has 16 protons, silicon has 14 protons, and gold has 79 protons.

Since an atom of one element can be distinguished from an atom of another element by the number of

protons in its nucleus, scientists are always interested in this number, and how this number differs between different elements. Therefore, scientists give this number a special name. An element's **atomic number** 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. The atomic number for hydrogen is 1, because every hydrogen atom has 1 proton. The atomic number for helium is 2 because every helium atom has 2 protons. What is the atomic number of carbon?

Of course, 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 an atomic number of 1. This means that an atom of hydrogen has one proton, and, if it's neutral, one electron as well. Gold, on the other hand, has an atomic number of 79, which means that an atom of gold has 79 protons, and, if it's neutral, and 79 electrons as well.

The **mass number** 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. The mass of an atom depends on the number of protons and neutrons. You have already learned 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). Counting the number of protons and neutrons tells scientists about the total mass of an atom.

$$\textit{mass number } A = (\textit{number of protons}) + (\textit{number of neutrons})$$

An atom's mass number is a very easy to calculate provided you know the number of protons and neutrons in an atom.

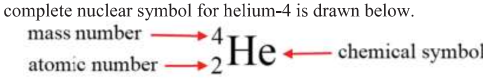
Example:
What is the mass number of an atom of helium that contains 2 neutrons?

Solution:
(number of protons) = 2 (Remember that an atom of helium always has 2 protons.)
(number of neutrons) = 2

mass number = (number of protons) + (number of neutrons)
mass number = 2 + 2 = 4

There are two main ways in which scientists frequently show the mass number of an atom they are interested in. It is important to note that the mass number is *not* given on the periodic table. These two ways include writing a nuclear symbol or by giving the name of the element with the mass number written.

To write a **nuclear symbol**, the mass number is placed at the upper left (superscript) of the chemical symbol and the atomic number is placed at the lower left (subscript) of the symbol. The complete nuclear symbol for helium-4 is drawn below.



The following nuclear symbols are for a nickel nucleus with 31 neutrons and a uranium nucleus with 146 neutrons.

$^{59}_{28}Ni$ $^{238}_{92}U$

In the nickel nucleus represented above, the atomic number 28 indicates the nucleus contains 28 protons, and therefore, it must contain 31neutrons in order to have a mass number of 59. The uranium nucleus has 92 protons as do all uranium nuclei and this particular uranium nucleus has 146 neutrons.

The other way of representing these nuclei would be Nickel-59 and Uranium-238, where 59 and 238 are the mass numbers of the two atoms, respectively. Note that the mass numbers (not the number of neutrons) is given to the side of the name.

Isotopes

Unlike the number of protons, which is always the same in atoms of the same element, the number of neutrons can be different, even in atoms of the same element. Atoms of the same element, containing the same number of protons, but different numbers of neutrons are known as **isotopes**. Since the isotopes of any given element all contain the same number of protons, they have the same atomic number (for example, the atomic number of helium is always 2). 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:
a) What is the atomic number and the mass number of an isotope of lithium containing 3 neutrons. A lithium atom contains 3 protons in its nucleus.
b) What is the atomic number and the mass number of an isotope of lithium containing 4 neutrons. A lithium atom contains 3 protons in its nucleus.

Solution:
a) atomic number = (number of protons) = 3
(number of neutrons) = 3
mass number = (number of protons) + (number of neutrons)
mass number = 3 + 3 = 6

b) atomic number = (number of protons) = 3
(number of neutrons) = 4
mass number = (number of protons) + (number of neutrons)
mass number = 3 + 4 = 7

Notice that because the lithium atom always has 3 protons, the atomic number for lithium is always 3. The mass number, however, is 6 in the isotope with 3 neutrons, and 7 in 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 constant uniform mixtures 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.

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. For most of our purposes in chemistry, we will do the same thing and deal with the average mass of the atoms. Luckily, aside from having different masses, most other properties of different isotopes are similar.

We can use what we know about atomic number and mass number to find the number of protons, neutrons, and electrons in any given atom or isotope. Consider the following examples:

Example: How many protons, electrons, and neutrons are in an atom of $^{40}_{19}K$?

Solution:
Finding the number of protons is simple. The atomic number, # of protons, is listed in the bottom right corner. # protons = 19.
For all atoms with no charge, the number of electrons is equal to the number of protons. # electrons =19.
The mass number, 40, is the sum of the protons and the neutrons. To find the # of neutron, subtract the number of protons from the mass number. # neutrons = 40 – 19 = 21.

Example: How many protons, electrons, and neutrons in an atom of zinc-65?

Solution:
Finding the number of protons is simple. The atomic number, # of protons, is found on the periodic table. All zinc atoms have # protons = 30.
For all atoms with no charge, the number of electrons is equal to the number of protons. # electrons =30.
The mass number, 65, is the sum of the protons and the neutrons. To find the # of neutron, subtract the number of protons from the mass number. # neutrons = 65 – 30 = 35.

Lesson Summary

- Electrons are a type of subatomic particle with a negative charge.
- Protons are a type of subatomic particle with a positive charge. Protons are bound together in an atom's nucleus as a result of the strong nuclear force.
- 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 the strong nuclear force.
- Protons and neutrons have approximately the same mass, but they 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 is equal to the number of protons in the nuclei of any of its atoms.
- The mass number 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.

Vocabulary

- Neutron: a subatomic particle with no charge
- Atomic mass unit (amu): a unit of mass equal to one-twelfth the mass of a carbon-twelve atom
- Atomic number: the number of protons in the nucleus of an atom
- Mass number: the total number of protons and neutrons in the nucleus of an atom
- Isotopes: atoms of the same element that have the same number of protons but different numbers of neutrons

Further Reading / Supplemental Material

- Jeopardy Game: http://www.quia.com/cb/36842.html
- For a Bill Nye video on atoms, go to http://www.uen.org/dms/. Go to the k-12 library. Search for "Bill Nye atoms". (you can get the username and password from your teacher)

2.3: Review Questions

Label each of the following statements as true or false.

1) The nucleus of an atom contains all of the protons in the atom.
2) The nucleus of an atom contains all of the electrons in the atom.
3) Neutral atoms must contain the same number of neutrons as protons.
4) Neutral atoms must contain the same number of electrons as protons.

Match the subatomic property with its description.

Sub-Atomic Particle	Characteristics
5) electron	a. has a charge of +1
6) neutron	b. has a mass of approximately 1/1840 amu
7) proton	c. is neither attracted to, nor repelled from charged objects

Indicate whether each statement is true or false.

8) An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
9) A neutral atom with 4 protons must have 4 electrons.
10) An atom with 7 protons and 7 neutrons will have a mass number of 14.
11) An atom with 7 protons and 7 neutrons will have an atomic number of 14.
12) A neutral atom with 7 electrons and 7 neutrons will have an atomic number of 14.

Use the periodic table to find the symbol for the element with:

13) 44 electrons in a neutral atom
14) 30 protons
15) An atomic number of 36

In the table below, Column 1 contains data for 5 different elements. Column 2 contains data for the same 5 elements, however different isotopes of those elements. Match the atom in the first column to its isotope in the second column.

Original element	Isotope of the same element
16) an atom with 2 protons and 1 neutron	a. a C (carbon) atom with 6 neutrons
17) a Be (beryllium) atom with 5 neutrons	b. an atom with 2 protons and 2 neutrons
18) an atom with an atomic number of 6 and mass number of 13	c. an atom with an atomic number of 7 and a mass number of 15
19) an atom with 1 proton and a mass number of 1	d. an atom with an atomic number of 1 and 1 neutron
20) an atom with an atomic number of 7 and 7 neutrons	e. an atom with an atomic number of 4 and 6 neutrons

Write the nuclear symbol for each element described:

21) 32 neutrons in an atom with mass number of 58

22) An atom with 10 neutrons and 9 protons.

Indicate the number of protons, neutrons, and electrons in each of the following atoms:

23) ${}^{4}_{2}He$ 24) Sodium-23 25) ${}^{1}_{1}H$

26) Iron-55 27) ${}^{37}_{17}Cl$ 28) Boron-11

29) ${}^{238}_{92}U$ 30) Uranium-235

2.4: Atomic Mass

Objectives:

- Explain what is meant by the atomic mass of an element.
- Calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

Introduction

In chemistry we very rarely deal with only one isotope of an element. We use a mixture of the isotopes of an element in chemical reactions and other aspects of chemistry, because all of the isotopes of an element react in the same manner. That means that we rarely need to worry about the mass of a specific isotope, but instead we need to know the average mass of the atoms of an element. Using the masses of the different isotopes and how abundant each isotope is, we can find the average mass of the atoms of an element. The **atomic mass** 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.

Calculating Atomic Mass

You can calculate the atomic mass (or average mass) of an element provided you know the **relative abundances** (the fraction of an element that is a given isotope) the element's naturally occurring isotopes, and the masses of those different isotopes. We can calculate this by the following equation:

$$\textbf{\textit{Atomic mass}} = (\%_1)(mass_1) + (\%_2)(mass_2) + \ldots$$

Look carefully to see how this equation is used in the following examples.

> **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:** Boron has two isotopes. We will use the equation:
>
> $$\textit{Atomic mass} = (\%_1)(mass_1) + (\%_2)(mass_2) + \ldots$$
>
> Isotope 1: $\%_1=0.20$ (write all percentages as decimals), $mass_1=10$
> Isotope 2: $\%_2=0.80$, $mass_2=11$
>
> Substitute these into the equation, and we get:
>
> $$\text{Atomic mass} = (0.20)(10) + (0.80)(11)$$
> $$\text{Atomic mass} = 10.8 \text{ amu}$$
>
> 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, 90.92% of the atoms are Ne-20, which is an isotope of neon with 10 neutrons and a mass of 19.99 amu. Another 0.3% of the atoms are Ne-21, which is an isotope of neon with 11 neutrons and a mass of 20.99 amu. The final 8.85% 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:**
> Neon has three isotopes. We will use the equation:
>
> $$\textit{Atomic mass} = (\%_1)(mass_1) + (\%_2)(mass_2) + \ldots$$
>
> Isotope 1: $\%_1=0.9092$ (write all percentages as decimals), $mass_1=19.99$
> Isotope 2: $\%_2=0.003$, $mass_2=20.99$
> Isotope 3: $\%_3=0.0885$, $mass_3=21.99$
>
> Substitute these into the equation, and we get:
>
> $$\text{Atomic mass} = (0.9092)(19.99) + (0.003)(20.99) + (0.0885)(21.99)$$
> $$\text{Atomic mass} = 20.17 \text{ amu}$$
>
> The mass of an average neon atom is **20.17 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, which is what we calculated in example 5, and the atomic mass of neon (symbol Ne) is 20.18, which is what we calculated in example 6. Take time to notice that not all periodic tables have the atomic number above the element's symbol and the mass number 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 and will be a decimal number.

Lesson Summary

- 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 element's atomic mass.

Vocabulary

- Atomic mass: the weighted average of the masses of the isotopes of an element

2.4: Review Questions

1) Copper has two naturally occurring isotopes. 69.15% of copper atoms are Cu-63 and have a mass of 62.93amu. The other 30.85% of copper atoms are Cu-65and have a mass of 64.93amu. What is the atomic mass of copper?

2) Chlorine has two isotopes, Cl-35 and Cl-37. Their abundances are 75.53% and 24.47% respectively. Calculate the atomic mass of chlorine.

2.5: The Nature of Light

Objectives

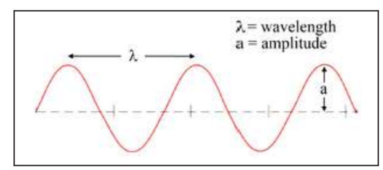
- When given two comparative colors or areas in the electromagnetic spectrum, identify which area has the higher wavelength, the higher frequency, and the higher energy.
- Describe the relationship between wavelength, frequency, and energy of light waves (EMR)

Introduction

Most of us are familiar with waves, whether they are waves of water in the ocean, waves made by wiggling the end of a rope, or waves made when a guitar string is plucked. Light, also called **electromagnetic radiation**, is a special type of energy that travels as a wave.

Light Energy

Before we talk about the different forms of light or electromagnetic radiation (EMR), it is important to understand some of the general characteristics that waves share.



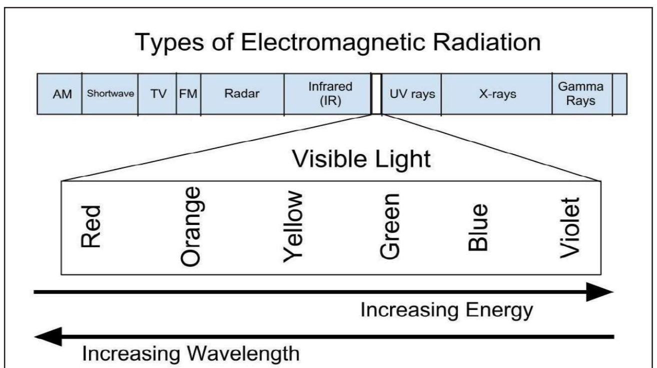
The high point of a wave is called the crest. The low point is called the trough. The distance from one point on a wave to the same point on the next wave is called the **wavelength** of the wave. You could determine the wavelength by measuring the distance from one trough to the next or from the top (crest) of one wave to the crest of the next wave. The symbol used for wavelength is the Greek letter lambda, λ.

Another important characteristic of waves is called frequency. The **frequency** of a wave is the number of waves that pass a given point each second. If we choose an exact position along the path of the wave and count how many waves pass the position each second, we would get a value for frequency. Frequency has the units of cycles/sec or waves/sec, but scientists usually just use units of 1/sec or Hertz (Hz).

All types of light (EMR) travels at the same speed, $3.00 \cdot 10^8$ m/s. Because of this, as the wavelength increases (the waves get longer), the frequency decreases (fewer waves pass). On the other hand, as the wavelength decreases (the waves get shorter), the frequency increases (more waves pass).

Electromagnetic waves (light waves) have an extremely wide range of wavelengths, frequencies, and energies. The **electromagnetic spectrum** is the range of all possible frequencies of electromagnetic radiation. The highest energy form of electromagnetic waves is gamma rays and the lowest energy form (that we have named) is radio waves.

On the far left of the figure above are the electromagnetic waves with the highest energy. These waves are called gamma rays and can be quite dangerous in large numbers to living systems. The next lowest energy form of electromagnetic waves is called x-rays. Most of you are familiar with the penetration abilities of these waves. They can also be dangerous to living systems. Next lower, in energy, are ultraviolet rays. These rays are part of sunshine and rays on the upper end of the ultraviolet range can cause sunburns and eventually skin



CC – Tracy Poulsen

cancer. The tiny section next in the spectrum is the visible range of light. These are the frequencies (energies) of the electromagnetic spectrum to which the human eye responds. The highest form of visible light energy is violet light, with red light having the lowest energy of all visible light. Even lower in the spectrum, too low in energy to see, are infrared rays and radio waves.

The light energies that are in the visible range are electromagnetic waves that cause the human eye to respond when those frequencies enter the eye. The eye sends signals to the brain and the individual "sees" various colors. The highest energy waves in the visible region cause the brain to see violet and as the energy of the waves decreases, the colors change to blue, green, then to yellow, orange, and red. When the energy of the wave is above or below the visible range, the eye does not respond to them. When the eye receives several different frequencies at the same time, the colors are blended by the brain. If all frequencies of light strike the eye together the brain sees white, and if there are no frequencies striking the eye the brain sees black.

All the objects that you see around you are light absorbers – that is, the chemicals on the surface of the objects absorb certain frequencies and not others. Your eyes detect the frequencies that strike your eye. Therefore, if your friend is wearing a red shirt, it means that the dye in that shirt absorbs every frequency except red and the red is reflected. When the red frequency from the shirt strikes your eye, your visual system sees red and you say the shirt is red. If your only light source was one exact frequency of blue light and you shined it on a shirt that absorbed every frequency of light except one exact frequency of red, then the shirt would look black to you because no light would be reflected to your eye. The light from many fluorescent types of light do not contain all the frequencies of sunlight and so clothes inside a store may appear to be a slightly different color than when you get them home.

Lesson Summary

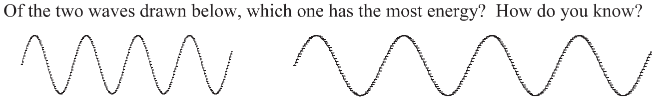
- Wave form energy is characterized by velocity, wavelength, and frequency.
- As the wavelength of a wave increases, its frequency decreases. Longer waves with lower frequencies have lower energy. Shorter waves with higher frequencies have higher energy.
- Electromagnetic radiation has a wide spectrum, including low energy radio waves and very high energy gamma rays.
- The different colors of light differ in their frequencies (or wavelengths).

Vocabulary

- Frequency of a wave: The number of waves passing a specific point each second.
- Wavelength: The distance between a point on one wave to the same point on the next wave (usually from crest to crest or trough to trough).
- Electromagnetic spectrum: A list of all the possible types of light in order of decreasing frequency, or increasing wavelength, or decreasing energy. The electromagnetic spectrum includes gamma rays, X-rays, UV rays, visible light, IR radiation, microwaves and radio waves.

2.5: Review Questions

1) Which color of visible light has the longer wavelength, red or blue?
2) What is the relationship between the energy of electromagnetic radiation and the frequency of that radiation?
3) Of the two waves drawn below, which one has the most energy? How do you know?



4) List the following parts of the electromagnetic spectrum in order of INCREASING energy: radio, gamma, UV, visible light, and infrared

5) List the visible colors of light in order of INCREASING energy.

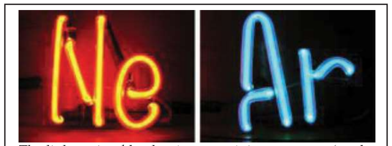
2.6: Atoms and Electromagnetic Spectra

Objectives

- Describe the appearance of an atomic emission spectrum.
- Explain that each element has a unique emission spectrum.
- Explain how an atomic (or emission) spectrum can be used to identify elements
- Describe an electron cloud that contains Bohr's energy levels.
- Explain the process through which an atomic spectrum is emitted according to Bohr's model of atoms.

Introduction

Electric light bulbs contain a very thin wire in them that emits light when heated. The wire is called a filament. The particular wire used in light bulbs is made of tungsten. A wire made of any metal would emit light under these circumstances but tungsten was chosen because the light it emits contains virtually every frequency and therefore, the light emitted by tungsten appears white. A wire made of some other element would emit light of some color that was not convenient for our uses. Every element emits light when energized by heating or passing electric current through it. Elements in solid form begin to glow when they are heated sufficiently and elements in gaseous form emit light when electricity passes through them. This is the source of light emitted by neon signs and is also the source of light in a fire.

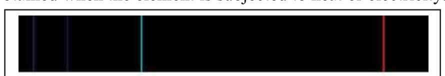


The light emitted by the sign containing neon gas (on the left) is different from the light emitted by the sign containing argon gas (on the right).

Each Element Has a Unique Spectrum

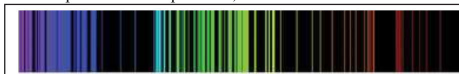
The light frequencies emitted by atoms are mixed together by our eyes so that we see a blended color. Several physicists, including Angstrom in 1868 and Balmer in 1875, passed the light from energized atoms through glass prisms in such a way that the light was spread out so they could see the individual frequencies that made up the light.

The **emission spectrum** (or **atomic spectrum**) of a chemical element is the unique pattern of light obtained when the element is subjected to heat or electricity.


Atomic spectrum of hydrogen

When hydrogen gas is placed into a tube and electric current passed through it, the color of emitted light is pink. But when the color is spread out, we see that the hydrogen

spectrum is composed of four individual frequencies. The pink color of the tube is the result of our eyes blending the four colors. Every atom has its own characteristic spectrum; no two atomic spectra are alike. The image below shows the emission spectrum of iron. Because each element has a unique emission spectrum, elements can be identified using them.

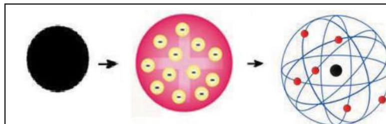

Atomic spectrum of iron.

You may have heard or read about scientists discussing what elements are present in the sun or some more distant star, and after hearing that, wondered how scientists could know what elements were present in a place no one has ever been. Scientists determine what elements are present in distant stars by analyzing the light that comes from stars and finding the atomic spectrum of elements in that light. If the exact four lines that compose hydrogen's atomic spectrum are present in the light emitted from the star, that element contains hydrogen.

Bohr's Model of the Atom

By 1913, the evolution of our concept of the atom had proceeded from Dalton's indivisible spheres idea to J. J. Thomson's plum pudding model and then to Rutherford's nuclear atom theory.

Rutherford, in addition to carrying out the brilliant experiment that demonstrated the presence of the atomic nucleus, also proposed that the electrons circled the nucleus in a planetary type motion. The solar system or planetary model of the atom was attractive to scientists because it was similar to something with which they were already familiar, namely the solar system.


Over time, our understanding of the atom has evolved. Dalton's model (on the left) was altered when Thomson discovered the electron and proposed the plum pudding model (in the middle). Rutherford discovered the nucleus and altered the model to the one on the right. Since then, Neils Bohr and other scientists discovered more about the location and energy of the electrons.

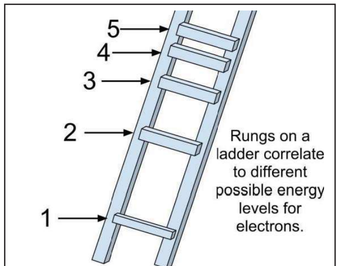
Unfortunately, there was a serious flaw in the planetary model. It was already known that when a charged particle (such as an electron) moves in a curved path, it gives off some form of light and loses energy in doing so. This is, after all, how we produce TV signals. If the electron circling the nucleus in an atom loses energy, it would necessarily have to move closer to the nucleus as it loses energy and would eventually crash into the nucleus. Furthermore, Rutherford's model was unable to describe how electrons give off light forming each element's unique atomic spectrum. These difficulties cast a shadow on the planetary model and indicated that, eventually, it would have to replaced.


Niels Bohr and Albert Einstein in 1925. Bohr received the Nobel prize for physics in 1922.

In 1913, the Danish physicist Niels Bohr proposed a model of the electron cloud of an atom in which electrons orbit the nucleus and were able to produce atomic spectrum. Understanding Bohr's model requires some knowledge of electromagnetic radiation (or light).

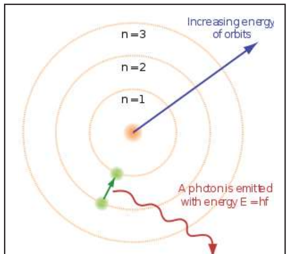
Energy Levels

Bohr's key idea in his model of the atom is that electrons occupy definite orbits that require the electron to have a specific amount of energy. In order for an electron to be in the electron cloud of an atom, it must be in one of the allowable orbits and it must have the precise energy required for that orbit. Orbits closer to the nucleus would require smaller amounts of energy for an electron and orbits farther from the nucleus would require the electrons to have a greater amount of energy. The possible orbits are known as **energy levels**. One of the weaknesses of Bohr's model was that he could not offer a reason why only certain energy levels or orbits were allowed.



Bohr proposed that electrons have specific locations (or energy levels) around the nucleus, much like there are specific steps on a ladder.
CC – Tracy Poulsen

Bohr hypothesized that the only way electrons could gain or lose energy would be to move from one energy level to another, thus gaining or losing precise amounts of energy. The energy levels are **quantized,** meaning that only specific amounts are possible. It would be like a ladder that had rungs only at certain heights. The only way you can be on that ladder is to be on one of the rungs and the only way you could move up or down would be to move to one of the other rungs. Suppose we had such a ladder with 10 rungs. Other rules for the ladder are that only one person can be on a rung and in normal state, the ladder occupants must be on the lowest rung available. If the ladder had five people on it, they would be on the lowest five rungs. In this situation, no person could move down because all the lower rungs are full. Bohr worked out rules for the maximum number of electrons that could be in each energy level in his model and required that an atom is in its normal state (ground state) had all electrons in the lowest energy levels available. Under these circumstances, no electron could lose energy because no electron could move down to a lower energy level. In this way, Bohr's model explained why electrons circling the nucleus did not emit energy and spiral into the nucleus.



In Bohr's model of the atom, electrons absorb energy to move to higher energy levels and release energy to move to lower energy levels.
Obtained from: http://en.wikipedia.org/wiki/File:Bohr_atom_model_English.svg

Bohr's Model and Atomic Spectra

The evidence used to support Bohr's model came from the atomic spectra. He suggested that an atomic spectrum is made by the electrons in an

atom moving energy levels. The electrons typically have the lowest energy possible, called **ground state**. If the electrons are given energy (through heat, electricity, light, etc) the electrons in an atom could absorb energy by jumping to a higher energy level or an **excited state**. The electrons then give off the energy in the form of a piece of light, called a **photon,** they had absorbed to fall back to a lower energy level. The energy emitted by electrons dropping back to lower energy levels would always be precise amounts of energy because the differences in energy levels were precise. This explains why you see specific lines of light when looking at an atomic spectrum – each line of light matches a specific "step down" that an electron can take in that atom. This also explains why each element produces a different atomic spectrum. Because each element has different acceptable energy levels for their electrons, the possible steps each element's electrons can take differ from all other elements.

Lesson Summary

- Bohr's model suggests each atom has a set of unchangeable energy levels and electrons in the electron cloud of that atom must be in one of those energy levels.
- Bohr's model suggests that the atomic spectra of atoms is produced by electrons gaining energy from some source, jumping up to a higher energy level, then immediately dropping back to a lower energy level and emitting the energy difference between the two energy levels.
- The existence of the atomic spectra is support for Bohr's model of the atom.
- Bohr's model was only successful in calculating energy levels for the hydrogen atom.

Vocabulary

- Emission spectrum (or atomic spectrum): The unique pattern of light given off by an element when it is given energy
- Energy levels: Possible orbits that an electron can have in the electron cloud of an atom.
- Ground state: to be in the lowest energy level possible
- Excited state: to be in a higher energy level
- Photon: a piece of electromagnetic radiation, or light, with a specific amount of energy
- Quantized: having specific amounts

Further Reading / Supplemental Links

- *A short discussion of atomic spectra and some animation showing the spectra of elements you chose and an animation of electrons changing orbits with the absorbtion and emission of light can be viewed at* Spectral Lines (http://www.colorado.edu/physics/2000/quantumzone/index.html)
- Tutorial: http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/linesp16.swf
- Tutorial: http://visionlearning.com/library/module_viewer.php?mid=51&l=&c3=
- Video: http://www.youtube.com/watch%3Fv%3DQI50GBUJ48s
- Fireworks & How Electrons Emit Photons video: http://www.youtube.com/watch%3Fv%3DncdmqhlTmGA

Review Questions

1) Bohr's model of the atom is frequently referred to as the "quantum model". Why? What does it mean to be quantized? How are electrons in atoms quantized?
2) Each element produces a unique pattern of light due to different energies within the atom. Why would this information be useful in analyzing a material?
3) It was known that an undiscovered element (later named helium) was on the sun before it was ever discovered on earth by looking at the sun's spectrum. How do scientists know that the sun contains helium atoms when no one has even taken a sample of material from the sun?
4) According to Bohr's theory, how can an electron gain or lose energy?
5) What happens when an electron in an excited atom returns to its ground level?
6) Why do electrons of an element release only a specific pattern of light? Why don't they produce all colors of light?
7) Use the following terms to explain how an electron releases a photon of light: electron, energy level, excited state, ground state, photon. Draw a picture if it is helpful.

2.6: Electron Arrangement in Atoms

Objectives

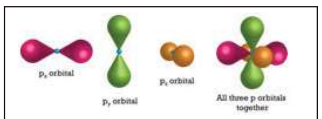
- List the order in which electron energy levels/sublevels will fill
- Write the electron configuration and abbreviated electron configuration for a given atom.

Introduction

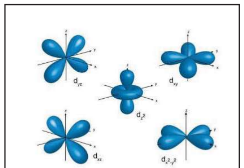
Chemists are particularly interested in the electrons in an atoms electron cloud. This is because the electrons determine the chemical properties of elements, such as what compounds the element will form and which reactions it will participate in. In this section, we will learn where the electrons are in atoms.

Electron Energy Levels

Although Bohr's model was particularly useful for hydrogen, it did not work well for elements larger than hydrogen. However, other physicists built on his model to create one that worked for all elements. It was found that the energy levels used for hydrogen were further composed of sublevels of different shapes. These sublevels were composed of orbitals in which the electrons were located.



The shape of p-orbitals


The shape of d-orbitals

The following table summarizes the possible energy levels and sublevels, including the number of orbitals that compose each sublevel and the number of electrons the sublevel can hold when completely filled.

Energy Level (related to the distance from the nucleus)	**Sublevel** (related to the shape)	**# of orbitals** in each sublevel	Maximum # of electrons possible
1	1s	1	2
2	2s	1	2
	2p	3	6
3	3s	1	2
	3p	3	6
	3d	5	10
4	4s	1	2
	4p	3	6
	4d	5	10
	4f	7	14
…	…	…	…

There are several patterns to notice when looking at the table of energy levels. Each energy level has one more sublevel than the level before it. Also, each new sublevel has two more orbitals. Can you predict what the 5^{th} energy level would look like?

When determining where the electrons in an atom are located, a couple of rules must be followed:

1. Each added electron enters the orbitals of the lowest energy available.
2. No more than two electrons can be placed in any orbital.

The Electron Configuration

It would be convenient if the sublevels filled in the order listed in the table, such as 1s, 2s, 2p, 3s, 3p, 3d, 4s, 4p, etc. However, this is not the order the electrons fill the sublevels. Remember, the electrons will always go to the lowest energy available. When that is taken into account, the actual filling order is:

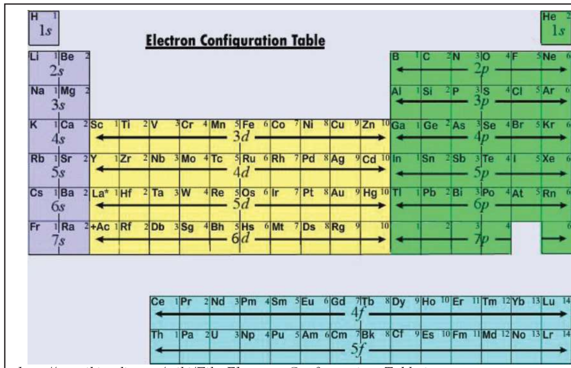
1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p…

Note that 4s has lower energy than 3d and, therefore, will fill first. The filling order gets more overlapped the higher you go.

An **electron configuration** lists the number of electrons in each used sublevel for an atom. For example, consider the element gallium, with 31 electrons. Its first two electrons would fit in the lowest energy possible, 1s. The next two would occupy 2s. 2p, with three orbitals, can hold its next 6 electrons. Gallium continues to fill up its orbitals, finally putting 1 electron in 4s. The electron configuration for gallium would be:

$$_{31}\text{Ga:}\ 1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^2\, 3d^{10}\, 4p^1$$

Although you can choose to memorize the list and how many electrons fit in each sublevel for the purpose of writing electron configurations, there is a way for us to find this order by simply using our periodic table.



http://en.wikipedia.org/wiki/File:Electron_Configuration_Table.jpg

Look at the different sections of the periodic table. You may have noticed that there are several natural sections of the periodic table. The first 2 columns on the left make up the first section; the six columns on the right make up the next section; the middle ten columns make up another section; finally the bottom fourteen columns compose the last section. Note the significance of these numbers: 2 electrons fit in any *s* sublevel, 6 electrons fit in any *p* sublevel, 10 electrons fit in any *d* sublevel, and 14 electrons fit in any *f* sublevel. The four sections described previously are known as the *s*, *p*, *d*, and *f* blocks respectively.

If you move across the rows starting at the top left of the periodic table and move across each successive row, you can generate the same order of filling orbitals that was listed before and also how many electrons total fit in each orbital. Starting at the top left, you are filling 1s. Moving onto the second row, 2s is filled followed by 2p. Continuing with the filling order you generate the list:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p…

An electron configuration lists the sublevels the electrons occupy and the number of electrons in each of those sublevels, written as superscripts.

Example: Write the electron configurations for
(a) potassium, K
(b) arsenic, As
(c) phosphorus, P

Solution:
(a) Potassium atoms have 10 protons and, therefore, 19 electrons. Using our chart, we see that the first sublevel to fill is *1s*, which can hold 2 of those 19 electrons. Next to fill is *2s*, which also holds 2 electrons. Then comes *2p* which holes 6. We keep filling up the sublevels until all 19 of the electrons have been placed in the lowest energy level possible. *4s* only has 1 electron in it, although it can hold up to 2 electrons, because there are only 19 electrons total in potassium. Its electron configuration is written as: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
(b) $_{33}$As: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$
(c) $_{15}$P: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

Abbreviated Electron Configuration

As the electron configurations become longer and longer, it becomes tedious to write them out. A shortcut has been devised so that writing the configurations is less tedious. The electron configuration for potassium is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$. The electron configuration for potassium is the same as the electron configuration for argon except that it has one more electron. The electron configuration for argon is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ and in order to write the electron configuration for potassium, we need to add only $4s^1$. It is acceptable to use [Ar] to represent the electron configuration for argon and $[Ar]4s^1$ to represent the electron configuration for potassium. Using this shortcut, the abbreviated electron configuration for calcium would be $[Ar]\ 4s^2$ and the electron configuration for scandium would be $[Ar]4s^2\ 3d^1$.

Even though the periodic table was organized according to the chemical behavior of the elements, you can now see that the shape and design of the table is a perfect reflection of the electron configuration of the atoms. This is because the chemical behavior of the elements is also caused by the electron configuration of the atoms.

Example: Write the abbreviated electron configurations for
(a) potassium, K
(b) arsenic, As
(c) phosphorus, P

Solution:
(a) $_{19}$K: $[Ar]\ 4s^1$
(b) $_{33}$As: $[Ar]\ 4s^2\ 3d^{10}\ 4p^3$
(c) $_{15}$P: $[Ne]\ 3s^2\ 3p^3$

Lesson Summary
- Electrons are located in orbitals, in various sublevels and energy levels of atoms
- Electrons will occupy the lowest energy level possible.
- It is possible to write the electron configuration of en element using a periodic table.

Vocabulary
- Electron configuration: a list that represents the arrangement of electrons of an atom.

Further Reading / Supplemental Links
- http://www.ethbib.ethz.ch/exhibit/pauli/elektronenspin_e.html
- http://www.lorentz.leidenuniv.nl/history/spin/goudsmit.html
- http://en.wikipedia.org/wiki

2.7: Review Questions
1) Which principal energy level holds a maximum of eight electrons?
2) Which sub-energy level holds a maximum of six electrons?
3) Which sub-energy level holds a maximum of ten electrons?
4) If all the orbitals in the first two principal energy levels are filled, how many electrons are required?
5) In which energy level and sub-level of the carbon atom is the outermost electron located?
6) How many electrons are in the 2p sub-energy level of a neutral nitrogen atom?

7) Which element's neutral atoms will have the electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^1$?
8) What energy level and sub-level immediately follow 5s in the filling order?
9) What is the outermost energy level and sub-level used in the electron configuration of potassium?

Write electron configurations for each of the following neutral atoms:

10) Magnesium
11) Nitrogen
12) Yttrium
13) Tin
14) Krypton
15) Cesium
16) Uranium

Write the abbreviated electron configuration for each of the following neutral atoms:

17) Fluorine
18) Aluminum
19) Titanium
20) Arsenic
21) Rubidium
22) Carbon

Chapter 3: The Organization of the Elements

3.1: Mendeleev's Periodic Table

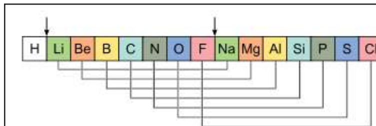
Objectives

- Describe the method Mendeleev used to make his periodic table.
- List the advantages and disadvantages Mendeleev's table had over other methods of organizing the elements.
- Explain how our current periodic table differs from Mendeleev's original table.

Introduction

During the 1800s, when most of the elements were being discovered, many chemists tried to classify the elements according to their similarities. In 1829, Johann Döbereiner noted chemical similarities in several groups of three elements and placed these elements into what he called triads. His groupings included the triads of 1) chlorine, bromine, and iodine, 2) sulfur, selenium, and tellurium, 3) calcium, strontium, and barium, and 4) lithium, sodium, and potassium. In all of the triads, the atomic weight of the second element was almost exactly the average of the atomic weights of the first and third element.

In 1864, John Newlands saw a connection between the chemical properties of elements and their atomic masses. He stated that if the known elements, beginning with lithium are arranged in order of increasing mass, the eighth element will have properties similar to the first, the ninth similar to the second, the tenth similar to the third, and so on. Newlands called his relationship the law of octaves, comparing the elements to the notes in a musical scale. Newlands tried to force all the known elements to fit into his octaves but many of the heavier elements, when discovered, did not fit into his patterns.

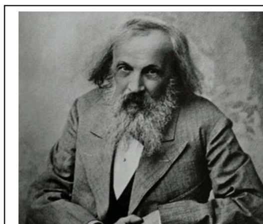


John Newlands' law of octaves suggested that, if elements are aligned in order of increasing mass, every eighth element would have similar properties.

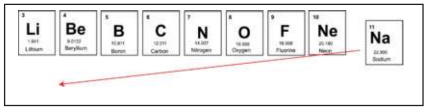
Mendeleev Organized His Table According to Chemical Behavior

By 1869, a total of 63 elements had been discovered. As the number of known elements grew, scientists began to recognize patterns in the way chemicals reacted and began to devise ways to classify the elements. Dmitri Mendeleev, a Siberian-born Russian chemist, was the first scientist to make a periodic table much like the one we use today.

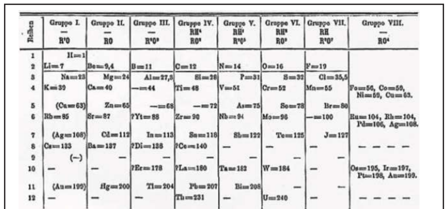
Mendeleev's table listed the elements in order of increasing atomic mass. Then he placed elements underneath other elements with similar chemical behavior. For example, lithium is a shiny metal, soft enough to be cut with a spoon. It reacts readily with oxygen and reacts violently with water.


Dmitri Mendeleev created the first periodic table in 1869.

When it reacts with water, it produces hydrogen gas and lithium hydroxide. As we proceed through the elements with increasing mass, we will come to the element sodium. Sodium is a shiny metal, soft enough to be cut with a spoon. It reacts readily with oxygen and reacts violently with water. When it reacts with water, it produces hydrogen gas and sodium hydroxide. You should note that the description of the chemical behavior of sodium is very similar to the chemical description of lithium. When Mendeleev found an element whose chemistry was very similar to a previous element, he placed it below the similar element.



Mendeleev avoided Newlands' mistake of trying to force elements into groups where their chemistry did not match, but still ran into a few problems as he constructed his table. Periodically, the atomic mass of elements would not be in the right order to put them in the correct group. For example, look at iodine and tellurium on your periodic table. Tellurium is heavier than iodine, but he put it before iodine in his table, because iodine has properties most similar to fluorine, chlorine, and bromine. Additionally, tellurium has properties more similar to the group with oxygen in it. On his table, he listed the element its place according to its properties and put a question mark (?) next to the symbol. The question mark indicated that he was unsure if the mass had been measured correctly.



Mendeleev's 1869 periodic table

Another problem Mendeleev encountered was that sometimes the next heaviest element in his list did not fit the properties of the next available place on the table. He would skip places on the table, leaving holes, in order to put the element in a group with elements with similar properties. For example, at the time the elements Gallium and Germanium had not yet been discovered. After zinc, arsenic was the next heaviest element he knew about, but arsenic had properties most similar to nitrogen and phosphorus, not boron. He left two holes in his table for what he claimed were undiscovered elements. Note the dashes (-) with a mass listed after it in his original table. These indicate places in which he predicted elements would later be discovered to fit and his predicted mass for these elements.

Mendeleev went further with his missing elements by predicting the properties of elements in those spaces. In 1871 Mendeleev predicted the existence of a yet-undiscovered element he called eka-aluminium (because its location was directly under aluminum's on the table). The table below compares the qualities of the element predicted by Mendeleev with actual characteristics of Gallium (discovered in 1875).

Mendeleev made similar predictions for an element to fit in the place next to silicon. Germanium, isolated in 1882, provided the best confirmation of the theory up to that time, due to its contrasting more clearly with its neighboring elements than the two previously confirmed predictions of Mendeleev do with theirs.

Property	Mendeleev's prediction for Eka-aluminium	Actual properties of Gallium
atomic mass	68	69.72
density (g/cm³)	6.0	5.904
melting point (°C)	Low	29.78
oxide's formula	Ea_2O_3 (density - 5.5 g cm^{-3}) (soluble in both alkalis and acids)	Ga_2O_3 (density - 5.88 g cm^{-3}) (soluble in both alkalis and acids)
chloride's formula	Ea_2Cl_6 (volatile)	Ga_2Cl_6 (volatile)

How was Mendeleev able to make such accurate predictions? He understood the patterns that appeared between elements within a family, as well as patterns according to increasing mass, that he was able to fill in the missing pieces of the patterns. The ability to make accurate predictions is was put Mendeleev's table apart from other organization systems that were made at the same time and is what led to scientists accepting his table and periodic law.

Property	Mendeleev's predictions for Eka-silicon	Actual properties of Germanium
atomic mass	72	72.61
density (g/cm³)	5.5	5.35
melting point (°C)	high	947
color	grey	grey
oxide type	refractory dioxide	refractory dioxide
oxide density (g/cm³)	4.7	4.7
oxide activity	feebly basic	feebly basic
chloride boiling point	under 100°C	86°C ($GeCl_4$)
chloride density (g/cm³)	1.9	1.9

Changes to our Modern Periodic Table

The periodic table we use today is similar to the one developed by Mendeleev, but is not exactly the same. There are some important distinctions:

Mendeleev's table did not include any of the noble gases, which were discovered later. These were added by Sir William Ramsay as Group 0, without any disturbance to the basic concept of the periodic table. (These elements were later moved to form group 18 or 8A.) Other elements were also discovered and put into their places on the periodic table.

As previously noted, Mendeleev organized elements in order of increasing atomic mass, with some problems in the order of masses. In 1914 Henry Moseley found a relationship between an element's X-ray wavelength and its atomic number, and therefore organized the table by nuclear charge (or atomic number) rather than atomic weight. Thus Moseley placed argon (atomic number 18) before potassium (atomic number 19) based on their X-ray wavelengths, despite the fact that argon has a greater atomic weight (39.9) than potassium (39.1). The new order agrees with the chemical properties of these elements, since argon is a noble gas and potassium an alkali metal. Similarly, Moseley placed cobalt before nickel, and was able to explain that tellurium should be placed before iodine, not because of an error in measuring the mass of the elements (as Mendeleev suggested), but because tellurium had a lower atomic number than iodine.

Moseley's research also showed that there were gaps in his table at atomic numbers 43 and 61 which are now known to be Technetium and Promethium, respectively, both radioactive and not naturally occurring. Following in the footsteps of Dmitri Mendeleev, Henry Moseley also predicted new elements.

You already saw that the elements in vertical columns are related to each other by their electron configuration, but remember that Mendeleev did not know anything about electron configuration. He placed the elements in their positions according to their chemical behavior. Thus, the vertical columns in Mendeleev's table were composed of elements with similar chemistry. These vertical columns are called **groups** or **families** of elements.

Lesson Summary

- The periodic table in its present form was organized by Dmitri Mendeleev.
- Mendeleev organized the elements in order of increasing atomic mass and in groups of similar chemical behavior. He also left holes for missing elements and used the patterns of his table to make predictions of properties of these undiscovered elements.
- The modern periodic table now arranges elements in order of increasing atomic number. Additionally, more groups and elements have been added as they have been discovered.

Vocabulary

- Periodic table: a tabular arrangement of the chemical elements according to atomic number.
- Mendeleev: the Russian chemist credited with organizing the periodic table in the form we use today.
- Moseley: the chemist credited with finding that each element has a unique atomic number

Further Reading / Supplemental Links

- Tutorial: Vision Learning: The Periodic Table
  http://visionlearning.com/library/module_viewer.php?mid=52&l=&c3=
- How the Periodic Table Was Organized (YouTube):
  http://www.youtube.com/watch%3Fv%3DCdkpoQk2LDE
- For several videos and video clips describing the periodic table, go to http://www.uen.org/dms/. Go to the k-12 library. Search for "periodic table". (you can get the username and password from your teacher)

3.1: Review Questions

1) What general organization did Mendeleev use when he constructed his table?
2) How did Mendeleev's system differ from Newlands's system?
3) Did all elements discovered at the time of Mendeleev fit into this organization system? How would the discovery of new elements have affecting Mendeleev's arrangement of the elements?
4) Look at Mendeleev's predictions for Germanium (ekasilicon). How was Mendeleev able to make such accurate predictions?
5) What problems did Mendeleev have when arranging the elements according to his criteria? What did he do to fix his problems?
6) What discovery did Henry Moseley make that changed how we currently recognize the order of the elements on the periodic table?

7) List three ways in which our current periodic table differs from the one originally made by Mendeleev.

3.2: Metals, Nonmetals, and Metalloids

Objectives

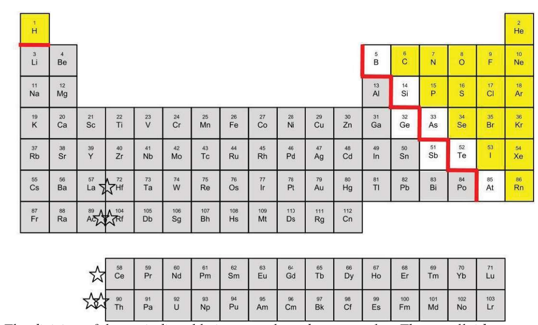
- Describe the differences among metals, nonmetals, and metalloids.
- Identify an element as a metal, nonmetal, or metalloid given a periodic table or its properties.

Introduction

In the periodic table, the elements are arranged according to similarities in their properties. The elements are listed in order of increasing atomic number as you read from left to right across a period and from top to bottom down a group. In this section you will learn the general behavior and trends within the periodic table that result from this arrangement in order to predict the properties of the elements.

Metals, Non-metals, and Metalloids

There is a progression from metals to non-metals across each row of elements in the periodic table. The diagonal line at the right side of the table separates the elements into two groups: the metals and the non-metals. The elements that are on the left of this line tend to be metals, while those to the right tend to be non-metals (with the exception of hydrogen which



The division of the periodic table into metals and non-metals. The metalloids are most of the elements along the line drawn. Additionally, the element hydrogen is a NONMETAL, even though it is on the left side of the periodic table.
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is a nonmetal). The elements that are directly on the diagonal line are metalloids, with some exceptions. Aluminum touches the line, but is considered a metal. Metallic character generally increases from top to bottom down a group and right to left across a period, meaning that francium (Fr) has the most metallic character of all of the discovered elements.

Most of the chemical elements are metals. Most metals have the common properties of being shiny, very dense, and having high melting points. Metals tend to be **ductile** (can be drawn out into thin wires) and **malleable** (can be hammered into thin sheets). Metals are good conductors of heat and electricity. All metals are solids at room temperature except for mercury. In chemical reactions, metals easily lose electrons to form positive ions. Examples of metals are silver, gold, and zinc.

Nonmetals are generally brittle, dull, have low melting points, and they are generally poor conductors of head heat and electricity. In chemical reactions, they ted to gain electrons to form negative ions. Examples of non-metals are hydrogen, carbon, and nitrogen.

Metalloids have properties of both metals and nonmetals. Metalloids can be shiny or dull. Electricity and heat can travel through metalloids, although not as easily as they can through metals. They are also called semimetals. They are typically semi-conductors, which means that they are elements that conduct electricity better than insulators, but not as well as conductors. They are valuable in the computer chip industry. Examples of metalloids are silicon and boron.

Lesson Summary

- There is a progression from metals to non-metals across each period of elements in the periodic table.
- Metallic character generally increases from top to bottom down a group and right to left across a period.

Vocabulary

- periodic law: states that the properties of the elements recur periodically as their atomic numbers increase
- ductile: can be drawn out into thin wires
- malleable: can be hammered into thin sheets

3.2: Review Questions

Label each of the following elements as a metal, nonmetal, or metalloid.

1) Carbon
2) Bromine
3) Oxygen
4) Plutonium
5) Potassium
6) Helium

Given each of the following properties, label the property of as that of a metal, nonmetal, or metalloid.

7) Lustrous
8) Semiconductors
9) Brittle
10) Malleable
11) Insulators
12) Conductors
13) Along the staircase

14) The elements mercury and bromine are both liquids at room temperature, but mercury is considered a metal and bromine is considered a nonmetal. How can that be? What properties do metals and nonmetals have?

3.3: Valence Electrons

Objectives

- Define valence electrons.
- Indicate the number of valence electrons for selected atoms.

Introduction

The electrons in the outermost shell are the **valence electrons** these are the electrons on an atom that can be gained or lost in a chemical reaction. Since filled *d* or *f* subshells are seldom disturbed in a chemical reaction, the valence electrons include those electrons in the outermost *s* and *p* sublevels.

Gallium has the following electron configuration.

Ga: $[Ar]\ 4s^2\ 3d^{10}\ 4p^1$

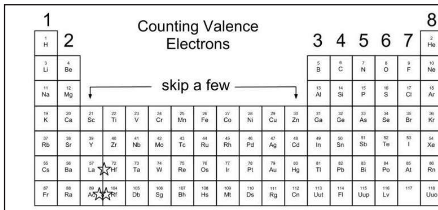
The electrons in the fourth energy level are further from the nucleus than the electrons in the third energy level. The *4s* and *4p* electrons can be lost in a chemical reaction, but not the electrons in the filled *3d* subshell. Gallium therefore has three valence electrons: the two in *4s* and one in *4p*.

Determining Valence Electrons

The number of valence electrons for an atom can be seen in the electron configuration. The electron configuration for magnesium is $1s^2\ 2s^2\ 2p^6\ 3s^2$. The outer energy level for this atom is $n=3$ and it has two electrons in this energy level. Therefore, magnesium has two valence electrons.

The electron configuration for sulfur is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$. The outer energy level in this atom is $n=3$ and it holds six electrons, so sulfur has six valence electrons.

The electron configuration for gallium is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$. The outer energy level for this atom is $n=4$ and it contains three electrons. You must recognize that even though the 3d sub-level is mixed in among the 4s and 4p sub-levels, 3d is NOT in the outer energy level and therefore, the electrons in the 3d sub-level are NOT valence electrons. Gallium has three electrons in the outer energy level and therefore, it has three valence electrons. The identification of valence electrons is vital because the chemical behavior of an element is determined primarily by the arrangement of the electrons in the valence shell.



The number of valence electrons in an atom can be easily found by counting the s and p columns in the periodic table.
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This pattern can be summarized very easily, by merely counting the *s* and *p* blocks of the periodic table to find the total number of valence electrons. One system of numbering the groups on the periodic table numbers the *s* and *p* block groups from 1A to 8A. The number indicates the number of valence electrons.

Lesson Summary
- Valence electrons are the electrons in the outermost principal quantum level of an atom.
- The number of valence electrons is important, because the chemical behavior of an element depends primarily by the arrangement of the electrons in the valence shell.

Vocabulary
- Valence electrons: the electrons in the outermost energy level of an atom.

3.3: Review Questions
1) How many valence electrons are present in the following electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^3$?
2) How many valence electrons are present in the following electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$?

For each of the following atoms, indicate the total number of valence electrons in each atom:
3) Fluorine
4) Bromine
5) Sodium
6) Cesium
7) Aluminum
8) Gallium
9) Argon
10) Krypton

3.4: Families and Periods of the Periodic Table
Objectives
- Give the name and location of specific groups on the periodic table, including alkali metals, alkaline earth metals, noble gases, halogens, and transition metals.
- Explain the relationship between the chemical behavior of families in the periodic table and their electron configuration.
- Identify elements that will have the most similar properties to a given element.

Introduction

The chemical behavior of atoms is controlled by their electron configuration. Since the families of elements were organized by their chemical behavior, it is predictable that the individual members of each chemical family will have similar electron configurations.

Families of the Periodic Table

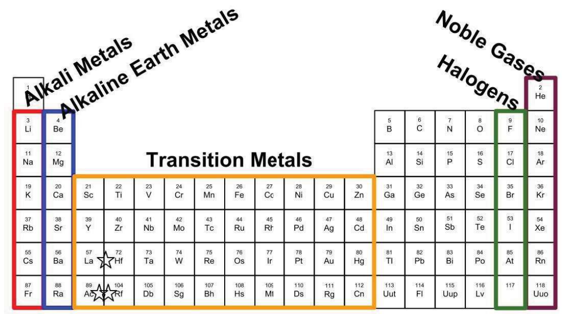
Remember that Mendeleev arranged the periodic table so that elements with the most similar properties were placed in the same group. A **group** is a vertical column of the periodic table. All of the 1A elements have one valence electron. This is what causes these elements to react in the same ways as the other members of the family. The elements in 1A are all very reactive and form compounds in the same ratios with similar properties with other elements. Because of their similarities in their chemical properties, Mendeleev put these elements into the same group. Group 1A is also known as the **alkali metals**. Although most metals tend to be very hard, these metals are actually soft and can be easily cut.

Group 2A is also called the **alkaline earth metals**. Once again, because of their similarities in electron configurations, these elements have similar properties to each other.

The same pattern is true of other groups on the periodic table. Remember, Mendeleev arranged the table so that elements with the most similar properties were in the same group on the periodic table.

It is important to recognize a couple of other important groups on the periodic table by their group name. Group 7A (or 17) elements are also called **halogens**. This group contains very reactive nonmetals elements.

The **noble gases** are in group 8A. These elements also have similar properties to each other, the most significant property being that they are extremely unreactive rarely forming compounds. We will learn the reason for this later, when we discuss how compounds form. The elements in this group are also gases at room temperature.



Families of the periodic table.
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An alternate numbering system numbers all of the *s*, *p*, and *d* block elements from 1-18. In this numbering system, group 1A is group 1; group 2A is group 2; the halogens (7A) are group 17; and the noble gases (8A) are group 18. You will come across periodic table with both numbering systems. It is important to recognize which numbering system is being used and to be able to find the number of valence electrons in the main block elements regardless of which numbering systems is being used.

Periods of the Periodic Table

If you can locate an element on the Periodic Table, you can use the element's position to figure out the energy level of the element's valence electrons. A **period** is a horizontal row of elements on the periodic

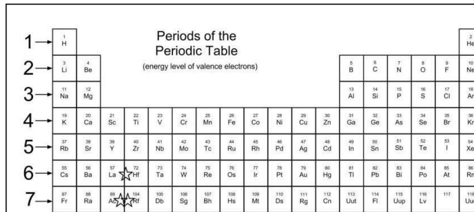

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table. For example, the elements sodium (Na) and magnesium (Mg) are both in period 3. The elements astatine (At) and radon (Rn) are both in period 6.

Lesson Summary
- The vertical columns on the periodic table are called groups or families because of their similar chemical behavior.
- All the members of a family of elements have the same number of valence electrons and similar chemical properties
- The horizontal rows on the periodic table are called periods.

Vocabulary
- Group (family): a vertical column in the periodic table
- Alkali metals: group 1A of the periodic table
- Alkali earth metals: group 2A of the periodic table
- Halogens: group 7A of the periodic table
- Noble gases: group 8A of the periodic table
- Transition elements: groups 3 to 12 of the periodic table

Further Reading / Supplemental Links
- http://www.wou.edu/las/physci/ch412/perhist.htm
- http://www.aip.org/history/curie/periodic.htm
- http://web.buddyproject.org/web017/web017/history.html
- http://www.dayah.com/periodic
- http://www.chemtutor.com/perich.htm

3.4: Review Questions
Multiple Choice
1) Which of the following elements is in the same family as fluorine?
   a) silicon
   b) antimony
   c) iodine
   d) arsenic
   e) None of these.
2) Elements in a _______________ have similar chemical properties.
   a) period
   b) family
   c) both A and B
   d) neither A nor B
3) Which of the following elements would you expect to be most similar to carbon?
   a) Nitrogen
   b) Boron
   c) Silicon

Give the name of the family in which each of the following elements is located:
4) astatine
5) krypton
6) barium
7) francium

Which family is characterized by each of the following descriptions?
8) A very reactive family of nonmetals
9) Have 7 valence electrons
10) A nonreactive family of nonmetals
11) Forms colorful compounds
12) Have 2 valence electrons
13) A very reactive family of metals

3.5: Periodic Trends
Objectives
- Explain what is meant by the term periodic law
- Describe the general trend in atomic size for groups and periods.
- Describe the trends that exist in the periodic table for ionization energy.
- Describe the trends that exist in the periodic table for electronegativity.

Introduction

We have talked in great detail about how the periodic table was developed, but we have yet to talk about where the periodic table gets its name. To be periodic means to "have repeating cycles" or repeating patterns. In the periodic table, there are a number of physical properties that are "trend-like". This means is that as you move down a group or across a period, you will see the properties changing in a general direction.

The periodic table is a powerful tool that provides a way for chemists to organize the chemical elements. The word "periodic" means happening or recurring at regular intervals. The periodic law states that the properties of the elements recur periodically as their atomic numbers increase. The electron configurations of the atoms vary periodically with their atomic number. Because the physical and chemical properties of elements depend on their electron configurations, many of the physical and chemical properties of the elements tend to repeat in a pattern.

The actual repeating trends that are observed have to do with three factors. These factors are:

(1) The number of protons in the nucleus (called the nuclear charge).
(2) The number of energy levels holding electrons (and the number of electrons in the outer energy level).
(3) The number of electrons held between the nucleus and its outermost electrons (called the **shielding effect**). This affects the attraction between the valence electrons and the protons in the nucleus.

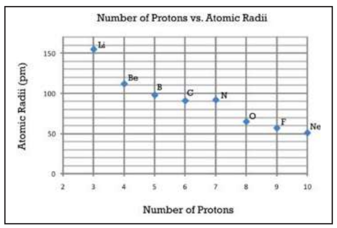
Trends in Atomic Radius

The atomic radius is a way of measuring the size of an atom. Although this is difficult to directly measure, we are, in essence, looking at the distance from the nucleus to the outermost electrons.

Let's look at the atomic radii or the size of the atom from the top of a family or group to the bottom. Take, for example, the Group 1 metals. Each atom in this family (and all other main group families) has the same number of electrons in the outer energy level as all the other atoms of that family. Each row (period) in the periodic table represents another added energy level. When we first learned about principal energy levels, we learned that each new energy level was larger than the one before. Energy level 2 is larger than energy level 1, energy level 3 is larger than energy level 2, and so on. Therefore, as we move down the periodic table from period to period, each successive period represents the addition of a larger energy level.

You can imagine that with the increase in the number of energy levels, the size of the atom must increase. The increase in the number of energy levels in the electron cloud takes up more space. Therefore the trend within a group or family on the periodic table is that *the atomic size increases with increased number of energy levels.*

In order to determine the trend for the periods, we need to look at the number of protons (nuclear charge), the number of energy levels, and the shielding effect. For a row in the periodic table, the atomic number still increases (as it did for the groups) and thus the number of protons would increase. When we examine the energy levels for period 2, we find that the outermost energy level does not change as we increase the number of electrons. In period 2, each additional electron goes into the second energy level. So the number of energy levels does not go up. As we move from left to right across a period, the number of electrons in the outer energy level increases but it is the same outer energy level.



Looking at the elements in period 2, the number of protons increases from lithium with three protons, to fluorine with nine protons. Therefore, the nuclear charge increases across a period. Meanwhile, the number of energy levels occupied by electrons remains the same. The numbers of electrons in the outermost energy level increases from left to right across a period, but how will this affect the radius? With an increase in nuclear charge, there is an increase in the pull between the protons and the outer level, pulling the outer electrons toward the nucleus. The net result is that *the atomic size decreases going across the row.*

Considering all the information about atomic size, you will recognize that the largest atom on the periodic table is all the way to the left and all the way to the bottom, francium, #87, and the smallest atom is all the way to the right and all the way to the top, helium, #2.

The fact that the atoms get larger as you move downward in a family is probably exactly what you expected before you even read this section, but the fact that the atoms get smaller as you move to the right across a period is most likely a big surprise. Make sure you understand this trend and the reasons for it.

Example: Which of the following has a greater radius?
(a) As or Sb
(b) Ca or K
(c) Polonium or Sulfur

Solution:
(a) Sb because it is below As in Group 15.
(b) K because it is further to the left on the periodic table.
(c) Polonium because it is below Sulfur in Group 16.

Periodic Trends in Ionization Energy

Lithium has an electron configuration of $1s^2\ 2s^1$. Lithium has one electron in its outermost energy level. In order to remove this electron, energy must be added. Look at the equation below:

$$Li(g) + \text{energy} \rightarrow Li^{+}(g) + e^{-}$$

With the addition of energy, a lithium ion can be formed from the lithium atom by losing one electron. This energy is known as the ionization energy. The **ionization energy** is the energy required to remove the most loosely held electron from a gaseous atom. The higher the value of the ionization energy, the harder it is to remove that electron.

Ionization Energies for some Group 1 Elements	
Element	**First Ionization Energy**
Lithium, Li	520 kJ/mol
Sodium, Na	495.5 kJ/mol
Potassium, K	801 kJ/mol

Ionization Energies for Period 2 Elements	
Element	**Ionization Energy**
Lithium, Li	520 kJ/mol
Beryllium, Be	899 kJ/mol
Boron, B	801 kJ/mol
Carbon, C	1086 kJ/mol
Nitrogen, N	1400 kJ/mol
Oxygen, O	1314 kJ/mol
Fluorine, F	1680 kJ/mol

Consider the ionization energies for the elements in group 1A of the periodic table, the alkali metals. Comparing the electron configurations of lithium to potassium, we know that the electron to be removed is further away from the nucleus, as the energy level of the valence electron increase. Because potassium's valence electron is further from the nucleus, there is less attraction between this electron and the protons and it requires less energy to remove this electron. As you move down a family (or group) on the periodic table, the ionization energy decreases.

We can see a trend when we look at the ionization energies for the elements in period 2. When we look closely at the data presented in the table above, we can see that as we move across the period from left to right, in general, the ionization energy increases. As we move across the period, the atoms become smaller which causes the nucleus to have greater attraction for the valence electrons. Therefore, as you move from left to right in a period on the periodic table, the ionization energy increases.

Example: Which of the following has a greater ionization energy?
(a) As or Sb
(b) Ca or K
(c) Polonium or Sulfur

Solution:
(a) As because it is above Sb in Group 15.
(b) Ca because it is further to the right on the periodic table.
(c) S because it is above Po in Group 16.

Periodic Trends in Electronegativity

Around 1935, the American chemist Linus Pauling developed a scale to describe the attraction an element has for electrons in a chemical bond. This is the **electronegativity**.