CHAPTER 11: MODERN ATOMIC THEORY

INTRODUCTION

It is difficult to get a mental image of atoms because we can't see them. Scientists have produced models which account for the behavior of atoms by making observations about the properties of atoms. So we know quite a bit about these tiny particles which make up matter even though we cannot see them. In this chapter you will learn how chemists believe atoms are structured.

GOALS FOR THIS CHAPTER

1. Learn how electromagnetic radiation is characterized, and what different types of electromagnetic radiation there are. (Section 11.1)
2. Know what happens when a hydrogen atom in an excited state loses energy. (Section 11.2)
3. Be able to cite the evidence for an excited atom possessing discrete quanta of energy. (Section 11.2)
4. Be able to explain what the Bohr model of the atom says about electron movement and know why the Bohr model is not correct. (Section 11.3)
5. Know what the wave mechanical model of the atom tells us about the location of an electron. (Section 11.4)
6. Know the physical shapes, the relative distances from the nucleus, and labels (symbols) for the orbitals of hydrogen through the third principal energy level. (Section 11.5)
7. Know what a probability map of an orbital represents. (Section 11.5)
8. Know that electron spin limits the number of electrons in each orbital to two. (Section 11.6)
9. Know how to write electron configurations and box diagrams for the elements, using the periodic table. (Sections 11.7 and 11.8)
10. Understand the relationship between the arrangement of atoms in the periodic table and the arrangement of electrons in orbitals. (Section 11.8)
11. Understand what atomic trends exist in the periodic table. (Section 11.9)

QUICK DEFINITIONS

Electromagnetic radiation — A form of energy called radiant energy that travels at the speed of light with wave-like behavior. (Section 11.1)

Wavelength — The distance from one wave peak to the next peak, abbreviated lambda, λ. (Section 11.1)

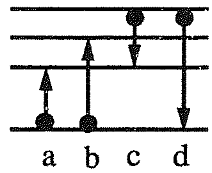
Frequency	Number of waves passing a particular point per second, abbreviated nu, ν. (Section 11.1)
Photons	Separate bundles of electromagnetic energy. (Section 11.1)
Excited state	The state of an electron in an atom with excess energy. (Section 11.2)
Ground state	An atom in its lowest energy state. (Section 11.2)
Quantized	The amount of energy associated with each energy level has a certain value. The energy amounts are fixed for each level. The fixed amount of energy is called a quantum of energy, so the energy values are said to be quantized. (Section 11.2)
Wave mechanical model	An atomic model which describes the distribution of electrons into orbitals. The wave mechanical model differs from the Bohr model, which suggested that electrons moved around the nucleus in circular orbits. (Section 11.4)
Orbital	A space described by the probability of finding an electron within that particular volume. (Section 11.5)
1*s* orbital	The orbital which is closest to the nucleus and of the lowest energy. (Section 11.5)
Principal energy levels	The major energy levels found in atoms. We refer to them as n=1, n=2, n=3, and so on. (Section 11.5)
Sublevels	Subdivisions of the principal energy levels. (Section 11.5)
Pauli Exclusion Principle	An atomic orbital can contain a maximum of two electrons, and they must be of opposite spin. (Section 11.6)
Electron configuration	The arrangement of electrons into principal energy levels and sublevels. We show the electron configuration by using symbols such as $2s^1$. The 2 refers to the principal energy level, the *s* to the type of sublevel, and the superscript 1 tells how many electrons occupy the orbital. The electron configuration for lithium is $1s^2 2s^1$. (Section 11.7)

Orbital diagram	Shows the arrangement of electrons using boxes to represent the orbitals. A number on top of the box tells the principal energy level, and a letter tells the type of sublevel. The number of arrows tells how many electrons are in the orbital. An orbital diagram is also called a box diagram. (Section 11.7)
Lanthanide series	A group of elements beginning with the first element after lanthanum (cerium) which fill their *4f* orbitals. The lanthanides are usually shown below the main body of the periodic table. (Section 11.7)
Actinide series	A group of elements beginning with the element after actinium (thorium), which fill their *5f* orbitals. The actinides are usually shown below the main body of the periodic table, and below the lanthanides. (Section 11.7)
Valence electrons	Electrons in the principal energy level furthest from the nucleus (the highest energy level). (Section 11.7)
Core electrons	All electrons which are not in the highest energy level. These electrons are all closer to the nucleus than the valence electrons. (Section 11.7)
Metal	An element with a lustrous appearance, the ability to change shape without breaking, and which conducts heat and electricity. (Section 11.9)
Nonmetal	An element which usually has none of the characteristics of metals given above. Some elements such as iodine and carbon have some but not all the characteristics of metals. (Section 11.9)
Ionization energy	The amount of energy required to remove an electron from a gaseous atom. (Section 11.9)

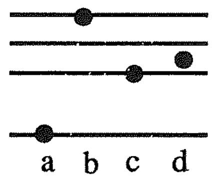
PRETEST

1. Which form of electromagnetic radiation has the longer wavelength, green light or red light?

2. Which arrow represents the path of an electron which is absorbing the greatest amount of energy?



3. Which electron is **not** in an allowed state?



4. The path in which an electron moves is not known exactly. Is this statement true or false?

5. How many orbitals are found in the 2s sublevel?

6. How many electrons can occupy the 3p sublevel?

7. If a 4s orbital contains two electrons, do they have the same spin, or opposite spin?

8. What is the electron configuration of chlorine, Cl?

9. What is the box notation for strontium, Sr?

10. Which element, Li, O, S, or Te has the greatest ionization energy?

PRETEST ANSWERS

1. Green light has a wavelength of around 5×10^{-7} m, while red light has a wavelength of around 7×10^{-7} m. Red light has the longer wavelength. (11.1)

2. Path a and b represent electrons that are absorbing energy. The electron in path b is absorbing a greater amount of energy than the electron in path a. (11.2)

3. Electron d is *between* allowed energy levels. It is not in an allowed state. (11.2)

4. The statement that the path of an electron is not known exactly is a true statement. (11.4)

5. The 2s sublevel contains one orbital. (11.5)

6. The 3p sublevel can contain a maximum of six electrons, two each in $3p_x$, $3p_y$, and $3p_x$. (11.5)

7. Two electrons that occupy a 4s sublevel orbital would be of opposite spin. (11.6)

8. The electron configuration of chlorine is

$1s^2 2s^2 2p^6 3s^2 3p^5$ (11.7)

9. The box notation for strontium is

1s	2s	2p	3s	3p	4s	3d
[↑↓]	[↑↓]	[↑↓][↑↓][↑↓]	[↑↓]	[↑↓][↑↓][↑↓]	[↑↓]	[↑↓][↑↓][↑↓][↑↓][↑↓]

4p	5s
[↑↓][↑↓][↑↓]	[↑↓]

(11.8)

10. Oxygen has a greater ionization energy than Li, S, or Te. (11.9)

CHAPTER REVIEW

11.1 ELECTROMAGNETIC RADIATION AND ENERGY

What Is Electromagnetic Radiation?

Electromagnetic radiation is a form of energy. This form of energy has some of the characteristics of waves and some of the characteristics of particles. A wave is characterized by wavelength, frequency, and speed. Wavelength is the distance between two wave peaks or troughs and is symbolized with the Greek letter lambda, λ. Frequency indicates how many waves pass per second and is symbolized with the Greek letter nu, ν. It is hard to imagine how energy can be both a wave and a particle, but experimental evidence justifies this view. This type of energy travels in packets called **photons**.

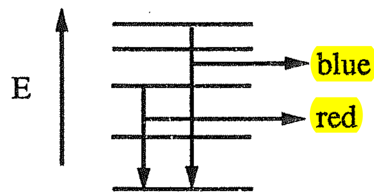
Height and weight. Determined by measuring tool.

11.2 THE ENERGY LEVELS OF HYDROGEN

How Does Electromagnetic Radiation Relate To the Study of Atoms?

All atoms can absorb and release extra energy. Atoms which have absorbed energy are in an **excited state**. Atoms in their lowest energy state are in the **ground state**. When atoms release energy, it is in the form of electromagnetic radiation. They emit a photon of light. The color of the light emitted is related to the amount of energy released. A photon of red light has less energy than a photon of blue light. The color of the photon and the amount of energy it contains depend upon the amount of energy the atom releases when it goes from an excited state to the ground state.

When an excited hydrogen atom that loses energy (goes from an excited state to a ground state) is observed to see what colors of light are emitted, it is found that only photons of certain colors are emitted. Since each color of photon is associated with a particular energy value, it means that hydrogen atoms are releasing only specific amounts of energy.



Hydrogen atoms always emit photons with these specific colors. That is, they always release the same amount of energy in the same packets. Because atoms jumping from a higher energy level to a lower one always release the same colored photons each time, their energy levels are quantized, that is, the energy of the photons is always one of the allowed values. We can tell this because photons of certain colors (certain energies) are the only ones we can detect.

11.3 THE BOHR MODEL OF THE ATOM

What Is the Bohr Model of the Atom?

Niels Bohr looked at the quantized energies produced when a hydrogen atom in an excited state loses energy. From these results, he proposed that the electron surrounding a hydrogen atom moves around the nucleus in circular orbits. When a hydrogen atom absorbed energy, the electron jumped from an orbit nearer the nucleus to an orbit farther away from the nucleus.

Likewise, when a hydrogen atom released energy, the electron jumped from an orbit farther away from the nucleus to one closer to the nucleus, releasing a photon of light of characteristic color.

Although Bohr's model explained why only light of certain colors is emitted when hydrogen atoms jump from an excited state to the ground state, it is basically incorrect. Electrons do not move in circular orbits.

Electrons would constantly emit energy by going in a circle.

11.4 THE WAVE MECHANICAL MODEL OF THE ATOM

What Is the Wave Mechanical Model of the Atom?

In thinking about how the electrons are distributed around an atom, two scientists, de Broglie and Schrödinger, theorized that since light can be thought of as both a particle and a wave, maybe an electron could also be considered as a particle and as a wave. When these ideas were treated mathematically, the results were the **wave mechanical model** of the atom. The wave mechanical model describes the behavior of the single electron of hydrogen, and the electrons of other atoms as well.

In the wave mechanical model, electrons are found in locations outside the nucleus called **orbitals**. Orbitals are not the same as circular orbits. Unlike our knowledge of planets in orbits, we cannot know at any one time exactly where an electron is. We can only know where an electron is likely to be. An electron is likely to be somewhere inside the volume of space described by the orbital.

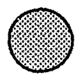
11.5 THE HYDROGEN ORBITALS

What Is Indicated By An Orbital Probability Map?

It can be hard to draw a picture of an orbital, because it represents an area of space where an electron is likely to be. We do not know exactly where the electron is. Theoretically, it is possible to find an electron very far away from the nucleus, even outside of its orbital. Chemists have decided to describe the shape of an orbital based on ninety percent probability. This means that 90 percent of the time, the electron will be found inside the volume of the orbital, and the other 10 percent, the electron will be found somewhere outside the orbital volume. By arbitrarily deciding where to stop including all the places an electron might be, everyone knows what is meant when someone draws an orbital shape. They mean that there is a 90 percent probability that the electron will be found inside that space.

How Do We Organize and Name the Orbitals Of Hydrogen?

When the electron of hydrogen is in the ground state, the electron occupies the **1*s* orbital**.



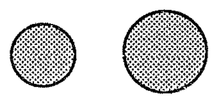
1*s*

The 1*s* orbital is the orbital closest to the nucleus, and describes a sphere. Remember that 90 percent of the time, the electron will be found somewhere within the sphere, and 10 percent of the time, the electron will be found somewhere outside the sphere.

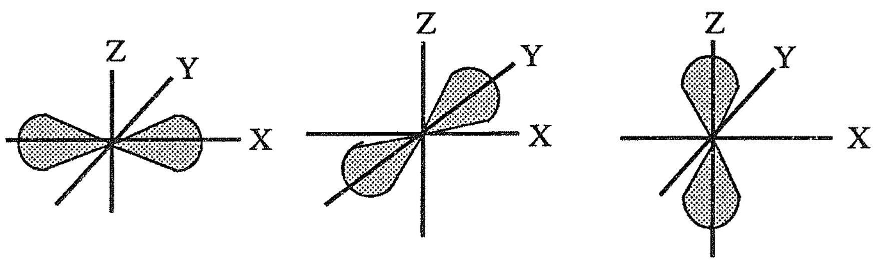
Hydrogen has other orbitals besides the 1*s*. You will recall that excited hydrogen atoms emit energy when electrons jump from a higher energy level to a lower energy level. Because the 1*s* orbital is the one of lowest energy (closest to the nucleus), there must be other orbitals of higher energy, farther away from the nucleus. Each of the energy levels is called a **principal energy level**. The first principal energy level is indicated with the number 1, the second with the number 2, the third with the number 3, and so on. As the number increases, the distance of the electrons from the nucleus increases. Electrons in the first principal energy level are closer to the nucleus than electrons in the third principal energy level.

Each of the principal energy levels is broken down into one or more sublevels. The first principal energy level has one sublevel. The second principal energy level has two sublevels. The third principal energy level has three sublevels, and so on. We use numbers and symbols to indicate which principal level and which sublevel an electron occupies. The 1*s* orbital of hydrogen means that the electron is in the first principal energy level (indicated by the number 1), and is in the *s* sublevel. The integer always indicates the principal level, and the letter always indicates the sublevel. The 1*s* sublevel has only one orbital, which is spherical in shape.

The second principal energy level has two sublevels, 2*s* and 2*p*. The 2*s* sublevel has one orbital, which is spherical in shape, like the 1*s* orbital. But the 2*s* orbital is larger than the 1*s*.



The 2*p* sublevel has three orbitals called $2p_x$, $2p_y$, and $2p_z$. The three 2*p* orbitals have a shape different from the 2*s* orbital. Each of the orbitals is dumbbell shaped, and is oriented along x, y or z axes. Remember that the pictures we draw of the shapes of orbitals tells us that 90 percent of the time, the electron will be found somewhere within that shape, and 10 percent of the time, somewhere outside the shape.



The third principal energy level has three sublevels, 3*s*, 3*p*, and 3*d*. The 3*s* sublevel has one orbital which is spherical in shape. In fact, the *s* sublevels of all the principal energy levels have orbitals which are spherical in shape. There are three 3*p* orbitals, $3p_x$, $3p_y$, and $3p_z$, each of which is dumbbell shaped along x, y or z axes. The 3*d* sublevel has five orbitals.

As you go from one principal energy level to the next, several facts about the orbitals emerge. All principal energy levels have an *s* sublevel, which contains one spherical orbital. The spherical orbital increases in size as you move farther away from the nucleus. Each principal energy level adds a new sublevel, with orbital shapes not found in the level below. For example, the second principal energy level has 2*p* orbitals, not found in the first energy level. The third principal energy level has 3*d* orbitals, not found in the second level.

11.6 THE WAVE MECHANICAL MODEL: FURTHER DEVELOPMENT

How Many Electrons Occupy Each Orbital?

Electrons are continuously spinning. When two electrons occupy the same orbital, it is necessary that they spin in opposite directions. Because of the need for opposite spin, only two electrons can occupy an orbital. This is called the **Pauli Exclusion Principle.**

Otherwise they would be the same electron.

How Can the Wave Mechanical Model Be Summarized?

There are seven principal components of the Wave Mechanical model, which are summarized below.

1. Atoms have a series of energy levels called principal energy levels which are described by whole numbers and symbolized by n. Level 1 corresponds to $n = 1$.
2. The energy increases as the value of n increases.
3. Each principal energy level contains one or more types of orbitals, called sublevels.
4. The number of sublevels present in a given principal energy level equals n.
5. The n value is always used to label the orbitals of a given principal level and is followed by a letter that indicates the type (shape) of the orbital.
6. An orbital can be empty or it can contain one or two electrons, but never more than two. If two electrons occupy the same orbital, they must have opposite spins.
7. The shape of an orbital does not indicate the details of electron movement. It indicates the probability of electron distribution for an electron residing in that orbital.

11.7 ELECTRON ARRANGEMENTS IN THE FIRST EIGHTEEN ATOMS IN THE PERIODIC TABLE

How Can We Show Which Orbitals Are Occupied By Electrons?

There are two ways that electrons are shown to be occupying orbitals. The electrons of lithium distributed into orbitals can be shown by $1s^2 2s^1$. This is called an **electron configuration** for lithium. $1s^2$ indicates that the s sublevel of the first principal energy level contains two electrons.

The superscript number tells the number of electrons in the sublevel. $2s^1$ indicates that the s sublevel of the second principal energy level contains only one electron.

The second way to indicate how the electrons of an atom are distributed is to use an **orbital diagram**, or a **box diagram.** A box diagram for lithium would look like

$1s$	$2s$
$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\ \,}$

Each box represents one orbital. Remember that an orbital can hold only two electrons, so each box can have a maximum of two arrows, each one indicating an electron with opposite spin.

How Do Atoms With Atomic Numbers Through 18 Fill the Available Orbitals?

In general, atoms fill sublevels closer to the nucleus before they fill sublevels farther away from the nucleus. Hydrogen has its only electron in the first principal energy level, closest to the nucleus. The first eighteen elements fill the first principal level first, followed by the 2*s* and the 2*p*, then the 3*s* and the 3*p*.

The element with atomic number sixteen is sulfur. Let's look at both an electron configuration and a box diagram for sulfur. $1s^2 2s^2 2p^6 3s^2 3p^4$ is the electron configuration for sulfur. Notice that electrons fill the innermost sublevels first. When writing electron configurations, it is common to lump the electrons occupying orbitals in a kind of sublevel together. For example, all of the electrons occupying the three 2*p* orbitals are indicated with one symbol, $2p^6$, instead of creating a separate symbol for the $2p_x$, $2p_y$, and $2p_z$ orbitals.

The box diagram for sulfur would be

1*s*	2*s*	2*p*			3*s*	3*p*		
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑

Notice that each individual 2*p* and 3*p* orbital has a separate box, in contrast to the electron configuration, which lumps all the electrons in a kind of sublevel together. Notice also that the 3*p* sublevel is not completely filled. There are only four electrons in that sublevel. When drawing box diagrams, put one electron in each sublevel box before putting two electrons in any box. For example, the $3p_x$, $3p_y$, and $3p_z$ orbitals would each receive one electron before the $3p_x$ orbital is filled with two electrons. We draw the arrows for $3p_y$ and $3p_z$ pointing in the same direction to indicate that they spin in the same direction.

What Are Valence and Core Electrons? Outermost s and p.

Valence electrons are the electrons found in the principal energy level farthest from the nucleus. For example, in sulfur, the valence electrons are the 3*s* and 3*p* electrons. **Core electrons** are the ones which are not valence electrons. Core electrons are inside the valence electrons and closer to the nucleus.

11.8 ELECTRON CONFIGURATIONS AND THE PERIODIC TABLE

How Do Electrons Fill Orbitals for Elements with Atomic Numbers Above 18?

Element eighteen, argon, has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6$. We might expect that the element with atomic number of nineteen, potassium, to begin filling the 3*d* sublevel. But when we look at the location of potassium in the periodic table, we see that it is a member of the alkali metals, along with lithium and sodium. Lithium has its last electron in the 2*s* sublevel, and sodium has its last electron in the 3*s* sublevel. Since groups of elements in the periodic table have similar properties and these properties are determined by the arrangement of valence electrons, it is likely that potassium has its last electron in the 4*s* sublevel, not the 3*d* sublevel. Experimental evidence has shown this to be true. The electron configuration of potassium is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$. After the 4*s* sublevel is filled by element twenty, calcium, the 3*d* orbitals are filled. In the periodic table, the elements which are filling 3*d* orbitals are the **transition metals**.

We can make some generalizations from the fourth period of the periodic table about how orbitals in elements in the rest of the table are filled. If a principal energy level has *d* orbitals, the *s* orbital of the next higher principal level (n + 1) always fills before the *d* orbitals of the previous principal energy level. For example, if the 4*p* orbitals are full, the 5*s* orbital is filled first, then the 4*d*.

After the element lanthanum is a series of elements called the **lanthanides**, usually shown below the main body of the periodic table, as in Figure 11.27 of your textbook. These elements are filling their 4*f* orbitals. Another series of elements which occur after the element actinium are called the **actinides**. These elements are filling their 5*f* orbitals. They are also shown below the body of the table for convenience.

It is possible to use the periodic table to tell which sublevels the outermost or valence electrons occupy. For example, elements in group 1, the alkali metals, have one electron in the outermost sublevel, an *s* sublevel. Elements in group 2, the alkaline earth metals, have two electrons in the outermost sublevel, an *s* sublevel. Elements in group 3 have three electrons in their outermost sublevel, a *p* sublevel, and so on. For the representative elements, the group number tells how many electrons are in the outermost sublevel.

11.9 ATOMIC PROPERTIES AND THE PERIODIC TABLE

What Are the Properties of Metals and Nonmetals?

Metals have a lustrous or shiny appearance. They can be beat into thin sheets and pulled into wires (are malleable), and they conduct heat and electricity. Nonmetals generally lack the properties of metals, although some nonmetals have some of the properties of metals. For example, solid iodine is lustrous.

The chemical properties of metals and nonmetals are important to chemists. Metals tend to lose electrons to form positive ions, while nonmetals tend to gain electrons to form negative ions.

How Are Metals and Nonmetals Grouped In the Periodic Table?

Metals are found in the left side and in the center of the periodic table. Most of the elements are metals. Some metals lose electrons easier than other metals. In general, metals at the bottom of the periodic table lose electrons easier than metals at the top. For example, barium loses an electron easier than beryllium does. The outermost electrons in barium are 6*s* electrons and are farther away from the nucleus than the 2*s* electrons of berellium, and so the electrons of barium are not held as tightly.

Nonmetals are found in the upper right corner of the periodic table. Compared to the number of metals, there are fewer nonmetals. Nonmetals vary in their ability to gain electrons from metals. The nonmetals with the greatest tendency to gain electrons are found in the upper righthand corner.

Are All Elements Either Metals Or Nonmetals?

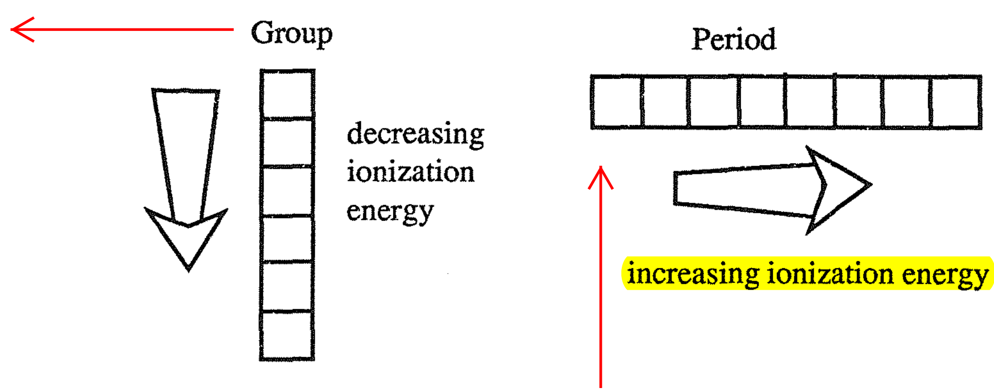
Some elements have properties of both the metals and the metalloids. These elements are located between the metals and nonmetals. Figure 11.31 in your textbook shows the divisions between metals, metalloids, and nonmetals.

How Does the Ionization Energy of An Atom Relate To Location In the Periodic Table?

The **ionization energy** of an atom is the amount of energy required to remove an electron from a gaseous atom. For any atom A

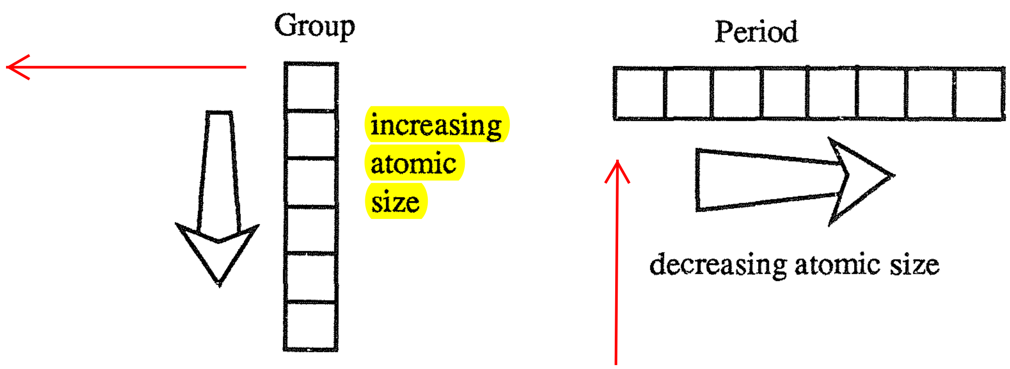
$$A(g) \xrightarrow[\text{energy}]{\text{ionization}} A^+(g) + e^-$$

Metals lose electrons easier than nonmetals, and metals at the bottom of the periodic table lose electrons easier than metals at the top of the periodic table. So ionization energy decreases from the top of the periodic table to the bottom and increases from the left side of the periodic table to the right.



How Does Atomic Size Relate to Location in the Periodic Table?

Atoms are not all the same size. Atoms get larger as we move from the top of the periodic table to the bottom, and they get smaller as we move from left to right.



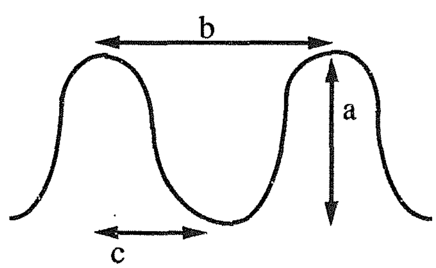
As we move from top to bottom, electrons are added to larger principal energy levels, so atoms become larger. As we move from right to left the number of electrons increases. However, all atoms in a period have their outermost electrons filling the same principal energy level. Because all the orbitals in a principal energy level would be expected to be the same size, we would expect that the size of an atom in a period would stay the same, not become smaller. As we move from left to right, the number of protons in the nucleus also increases. The increase in positive charge pulls the electrons closer, so as the number of protons increases from left to right, the closer the electrons are pulled, and the smaller the atomic size.

Example:

Which atom, silicon or lead, has a larger ionization energy and a smaller atomic size? Ionization energy decreases from top to bottom of the periodic table, so silicon would have a higher ionization energy than lead. Atomic size increases from top to bottom, so a silicon atom would be smaller than a lead atom.

LEARNING REVIEW

1. Which of the following represents the wavelength of electromagnetic radiation?



2. How does a microwave oven warm food?

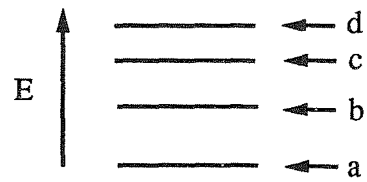
3. Which has the shorter wavelength, ultraviolet light or infrared light? See Figure 11.3 in your textbook for help with electromagnetic radiation.

4. Light can be thought of as waves of energy. There is also evidence that light exists in another form. What is the other form?

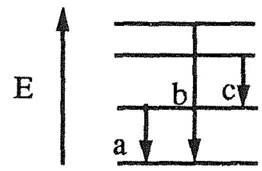
5. What is meant by the terms "ground state" and "excited state" of an atom?

6. A sample of helium atoms absorbs energy. Will the photons of light emitted by the helium atoms be found at all wavelengths? Explain your answer.

7. Which energy level represents the ground state?



8. Which quantum has the greatest energy?



9. What does the Bohr model of the atom say about electron movement?

10. What characteristics of light led de Broglie and Schrödinger to formulate a new model of the atom?

11. Which of the following statements about the wave mechanical model of the atom are true, and which are false?

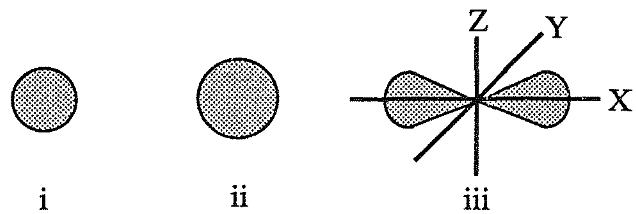
    a. The probability of finding an electron is the same in any location within an orbital.
    b. An electron will probably spend most of its time close to the nucleus.
    c. Electrons travel in circular orbits around the nucleus.

12. Which of the following statements about an orbital are true, and which are false?

    a. An electron will be found inside an orbital 90 percent of the time.
    b. An electron travels around the surface of an orbital.
    c. An electron cannot be found outside an orbital.

13. Consider the third principal energy level of hydrogen.

    a. How many sublevels are found in this level?
    b. How many orbitals are found in the 3*d* sublevel?
    c. Which shape represents a 3*p* orbital?



14. How many sublevels do you think would be found in the n=5 principal energy level?

15. What is meant by each part of the orbital symbol, $4p_x^{\,1}$?

16. How many electrons can occupy an orbital?

17. How does each column of the periodic table relate to electron configuration?

Use a periodic table such as the one found inside the front cover of your textbook to help answer questions 18, 19, and 20.

18. a. Write the complete electron configuration and the complete orbital diagram for aluminum.
    b. How many valence electrons and how many core electrons does aluminum have?

19. How many valence electrons are found in the elements beryllium, magnesium, calcium and strontium?

20. Write an electron configuration and box diagram for the elements below.

    a. vanadium
    b. copper
    c. bromine
    d. tin

21. How does each row of the periodic table relate to electron configuration?

22. What is characteristic of the electron configuration of the noble gases?

23. Which orbitals are filling in the lanthanide series elements?

24. Decide whether the elements below are representative elements or transition metals.

a.	Ar	c.	N
b.	Fe	d.	Sr

25. Which element in each pair would have a lower ionization energy?

a.	F and C	d.	Li and Rb
b.	O and As	e.	Ne and Rn
c.	Ca and Br	f.	Sr and Br

26. Which element in each pair would have the smaller atomic size?

a.	Ne and Xe	d.	F and Sr
b.	In and I	e.	Ba and Bi
c.	Na and Cs	f.	Cl and Al

ANSWERS TO LEARNING REVIEW

1. Wavelength is the distance from crest to crest, or trough to trough, so the correct answer is b.

2. The kind of electromagnetic radiation generated by the oven, microwave radiation, is of the right frequency to be absorbed by water in food. As the water molecules in food absorb microwave energy, their movement increases. The extra energy is transferred to other molecules in the food when they collide with the water molecules, so heat is transferred from rapidly moving water molecules to other molecules, causing the food to heat up.

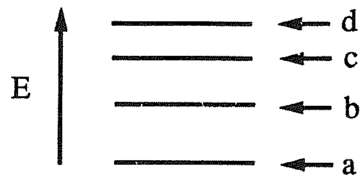
3. Infrared light has a wavelength of 10^{-4} meters and ultraviolet light has a wavelength of around 10^{-8} meters, so ultraviolet light has a shorter wavelength than infrared light.

4. There is evidence that light consists of packets of energy called photons.

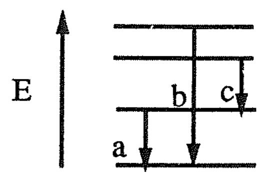
5. The lowest possible energy state of an atom is the ground state. When an atom has absorbed excess energy it is in an excited state.

6. When helium atoms absorb energy, some of the electrons move from the ground state to an excited state. When helium atoms lose this excess energy they will often emit light. The light is **not** of just any wavelength. Only certain wavelengths, corresponding to the differences in energy, are allowed.

7. The energy level marked a is the ground state because it is the level with the lowest amount of energy.



8. The quantum marked by b has the highest amount of energy because this excited state has more energy than the excited states below it.



9. Bohr's model said that electrons move in circular orbits around the nucleus. Each circular orbit represents an excited state farther from the nucleus.

10. The fact that light could act as both a wave and as a particle led de Broglie and Schrödinger to suggest that an electron might also exhibit characteristics of both waves and particles. When Schrödinger used these ideas to analyze the problem mathematically, the wave mechanical model of the atom was the result.

11. a. The probability of finding an electron within an orbital is 90 percent, but some locations within the orbital shape are more likely to contain an electron at any given time than are others. So the correct answer is false.
    b. The electron does tend to spend most of its time around the nucleus. This statement is true.
    c. Bohr thought that electrons traveled in circular orbits around the nucleus, but the current model says that electrons are found in orbitals. We do not know their exact paths. This statement is false.

12. a. The orbital shape represents a probability cloud. It is true that an electron will be found within the orbital 90 percent of the time.
    b. The orbital marks the area of 90 percent probability. It does not mark the surface on which the electron travels. This statement is false.
    c. Ten percent of the time, an electron will be found outside the orbital. This statement is false.

13. a. The third principal energy level, n=3, contains 3 sublevels.
    b. There are five orbitals in the 3*d* sublevel.
    c. i and ii represent *s* orbitals. iii is a *p* orbital.