CHAPTER 11

Modern Atomic Theory

INTRODUCTION

It is difficult to form a mental image of atoms because we can't see them. Scientists have produced models that 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.

CHAPTER DISCUSSION

This is yet another chapter in which models play a significant role. Recall that we like models to be simple, but we also need models to explain the questions we want to answer. For example, Dalton's model of the atom has the advantage of being quite simple and is useful when considering molecular-level views of solids, liquids, and gases and in representing chemical equations. However, it cannot explain fundamental questions such as why atoms "stick" together to form molecules. Scientists such as Thomson and Rutherford expanded the model of the atom to include subatomic particles. The model is more complicated than Dalton's, but it begins to explain chemical reactivity (electrons are involved) and formulas for ionic compounds. But questions still abound. For example, why are there similarities in reactivities of elements (periodic trends)?

To begin our understanding of modern atomic theory, let's first discuss some observations. For example, you have undoubtedly seen a fireworks display, either in person or on television. Where do the colors come from? How do we get so many different colors? It turns out that different salts, when heated, give off different colors. For example, copper(II) chloride is a characteristic green color, and lithium chloride is red. But why?

To answer this, let's consider a light bulb and why we see white light from it. The light that we see is part of a spectrum of electromagnetic radiation, which includes x-rays, UV rays, visible light, microwaves, etc. (see Section 11.2 in your text). A light bulb lights up because a thin filament in the bulb is heated (electric current is sent through the thin filament), and this heat (energy) is released as light. Realize that energy is related to the wavelength of electromagnetic radiation, and the wavelength (if visible) is related to color. Because the wire in the light bulb is heated, energy of all wavelengths is emitted, and therefore wavelengths of all colors are emitted. When all the colors of visible light are mixed, the result is white light.

So what does this mean about our fireworks? We know that electrons and protons attract each other; thus the electrons "want" to be close to the nucleus. We will include energy levels in our model of the structure of the atom. Light is given off because when the salts are heated, the electrons are excited and move to a higher energy state. When the sample is heated, these electrons are moved farther from the nucleus, but will go back to their original state. When the electrons return to their original (ground) state, energy is released, sometimes in the form of visible light. It is quite significant that the light is not white but has a characteristic color associated with each different salt. This must mean that not all wavelengths of light are present when a metallic salt (such as copper(II) chloride) is heated. In other words, the electrons cannot go to any excited state and then return to any other state. If this occurred, all wavelengths of light would be emitted, and the light would be white. It must mean that the electrons can go to only certain energy levels. This would account for only certain colors being emitted.

This is a surprising and non-intuitive result. Energy levels in an atom are "quantized;" that is, only certain levels are possible. This means, for example, that electrons can exist at one level or another level, but not in between the levels. The first person who tried to make sense of this idea was Neils Bohr who assumed a planetary-type model in which the electrons would orbit the nucleus much like we envision the planets orbiting the sun. That is, the electrons had known, predictable pathways. It is important to realize that this model is fundamentally incorrect. It assumes we know where the electrons are, and we know where they are going. The model we accept today just is not that simple. It turns out that we simply do not know how the electron travels.

The next step in trying to make sense of atomic structure was to consider the electron as moving as a wave (like electromagnetic radiation). While the specifics of this are beyond a first-year chemistry course, realize that this has the following results: 1) waves have a probability function associated with them and 2) waves can add up and cancel each other, thus we can get regions of high and low probability (and even zero probability of finding an electron). The firefly analogy in your text (Section 11.6) is a good one to read and understand in terms of how to think about electrons in an atom. Realize that the "size" or radius of an atom is arbitrary and that we usually consider that there is a 90% probability of finding an electron in the orbital (an unfortunate term that does not mean to suggest that the electron orbits the nucleus).

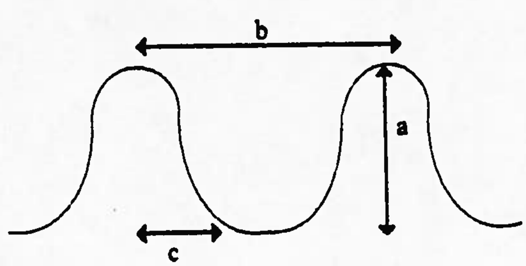
Imagine the firefly analogy but with more fireflies. The spherical pattern of one firefly makes sense. But what if there were more fireflies and (like electrons) they repelled each other? What would the patterns look like? These patterns would be impossible for us to predict. These, though, are analogous to the other orbitals such as the *p* and *d* orbitals. The shapes of these are not intuitive but have been mathematically determined. It is not expected that you look at the shape of a *p* orbital, for example, and say, "Of course that's the probability region of finding the electrons!" You would have no way of predicting this exact shape. Realize though that, as we could expect, the more electrons there are in an energy level, the more complex and large the orbitals become. Realize as well that orbitals are not physical things" but regions of probability.

The periodic table is considered in more detail in this chapter. Make sure to use it as a resource, not look at it as just another thing to memorize. The text provides an excellent discussion of this, and you should be able to answer the following:

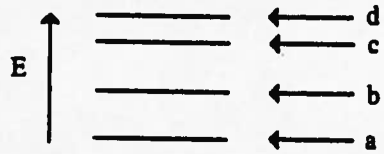
1. How are electron configurations consistent with the placement on the table? You don't have to memorize all configurations and should be able to quickly tell the configuration from the table.
2. In Chapter 4 we saw <u>what</u> the most stable charges were for many ions. Now you should be able to explain <u>why</u> they are the most stable charges.
3. Explain (don't just state) the trends of atomic radii and ionization energy, and explain how they are related to one another.

LEARNING REVIEW

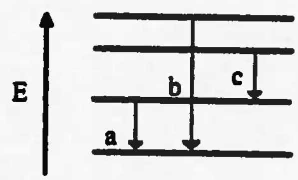
1. Which of the following represents the wavelength of electromagnetic radiation?



2. How does a microwave oven warm food?
3. Which has the shorter wavelength, ultraviolet light or infrared light? See Figure 11.4 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?

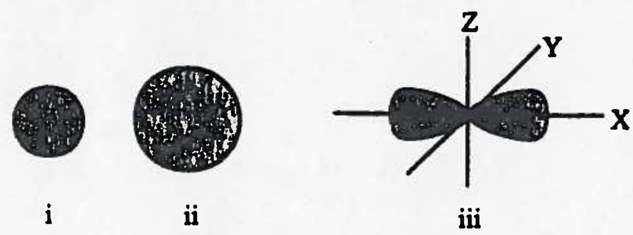


8. Which quantum has the greatest energy?



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% 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 $3p$ 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
    b. Fe
    c. N
    d. Sr
25. Which element in each pair would have a lower ionization energy?
    a. F and C
    b. O and As
    c. Ca and Br

d. Li and Rb

e. Ne and Rn

f. Sr and Br

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

a. Ne and Xe

b. In and I

c. Na and Cs

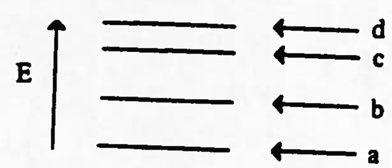
d. F and Sr

e. Ba and Bi

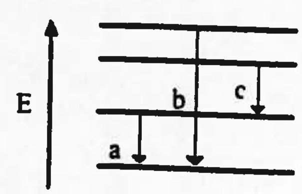
f. Cl and Al

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 and causes 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 both as 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%, but some locations within the orbital shape are more likely to contain an electron at any given time than are others. So this statement 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% of the time.
   b. The orbital marks the area of 90% 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 three sublevels.
   b. There are five orbitals in the $3d$ sublevel.
   c. i and ii represent *s* orbitals; iii is a *p* orbital.

14. Each principal energy level has n sublevels, so the fifth principal energy level, n=5, would have five sublevels.

15. 4 is the principal energy level, *p* is the sublevel type, x is the specific orbital within the *p* sublevel, and the superscript 1 says that there is one electron in the orbital.

16. Each orbital can hold two electrons.

17. The period or row number indicates which *s* and *p* orbitals are being filled for elements in that row. For example, antimony, which is in row five, has filled its $5s$ orbitals and $4d$ orbitals and is filling $5p$ orbitals.

18.
   a. $1s^2 2s^2 2p^6 3s^2 3p^1$

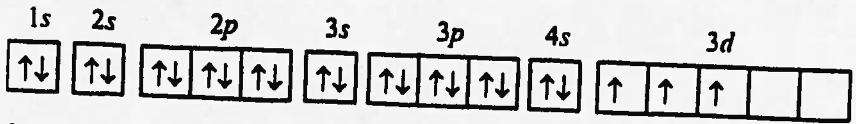
$1s$	$2s$	$2p$			$3s$	$3p$		
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑		

   b. Aluminum has three valence electrons and ten core electrons.

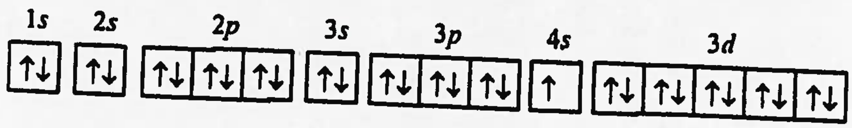
19. Each of these elements has two valence electrons.

20.

a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

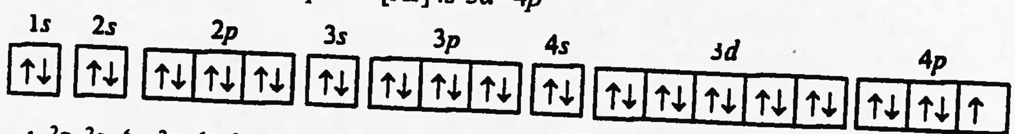


b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$ or $[Ar]4s^1 3d^{10}$

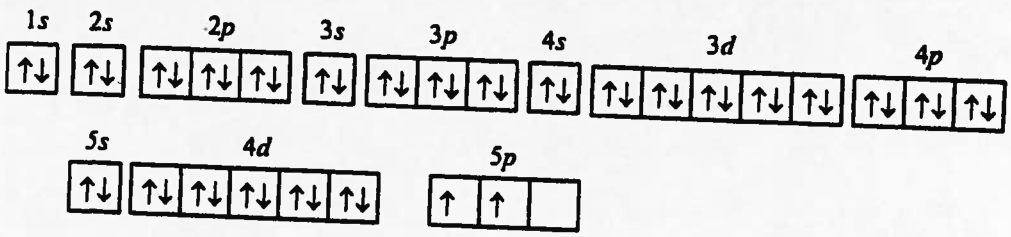


Note that copper does not completely fill $4s$ before filling $3d$.

c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$ or $[Ar]4s^2 3d^{10} 4p^5$



d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$



21. The group number on the top of each column of the periodic table is the same as the sum of $ns + np$ electrons for the highest principal energy level and is equal to the number of valence electrons for that element. For example, sulfur is in Group 6, and its electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^4$. The sum $ns + np$ for n=3 is six, which is the same as the group number. Sulfur has six valence electrons, $3s^2$ and $3p^4$.

22. The s and p orbitals for the principal energy level, which is the same as the row number, are full. That is, all the s and p orbitals contain a maximum of two electrons for a total of eight electrons.

23. In the lanthanide series elements, the $4f$ orbitals are filling.

24.

a. Representative element (noble gas)

b. Transition metal

c. Representative element

d. Representative element (alkaline earth metal)

25.

a. Carbon would have a lower ionization energy than fluorine.

b. Arsenic would have a lower ionization energy than oxygen.

c. Calcium would have a lower ionization energy than bromine.

d. Rubidium would have a lower ionization energy than lithium.

e. Radon would have a lower ionization energy than neon.

f. Strontium would have a lower ionization energy than bromine.

26.

a. Neon would have a smaller atomic size than xenon.

b. Iodine would have a smaller atomic size than indium.

c. Sodium would have a smaller atomic size than cesium.

d. Fluorine would have a smaller atomic size than strontium.

e. Bismuth would have a smaller atomic size than barium.

f. Chlorine would have a smaller atomic size than aluminum.