CHAPTER 3. INSIDE THE ATOM

What Is an Atom? A Closer View

The atomic theory was a powerful concept, a major step forward in understanding the nature of matter. Knowing that all substances are made up of atoms, and that the elements, or different kinds of atoms, had characteristic masses, enabled scientists to explain the identities of substances and helped them to understand how to change one substance into another. But just what were these unimaginably small atoms like? Why did different kinds of atoms behave so differently?

The Roman philosopher Lucretius thought that some atoms were smooth and round, but others had hooks on their surfaces, holding them together more tightly to produce hard, tightly bound substances like diamond and iron. Irritating substances or those with "sharp" odors must be composed of atoms that were at least partially hooked if they were so grating on the surfaces of the human body. These earliest speculations about the structure of atoms had no basis in experiment, of course. For hundreds of years little progress was made in understanding the nature of atoms, beyond the principle of atomic mass. Then around the beginning of the twentieth century a series of experiments prompted a major breakthrough in the understanding of the atom. In a burst of discovery marked by great intellectual excitement and creativity, the scientific world began to learn about the structure of the atom itself.

The Role of the Mysterious Rays

Mysterious rays made possible these critical discoveries. Invisible rays developed photographic film and could pass through walls and even the human body. Other rays emanated a ghostly glow and challenged accepted ideas about the nature of matter and energy. In a period of scientific discovery that gave us some of the most fascinating stories in the history of human thought, men and women worked almost obsessively to use these rays to unlock the secrets of matter.

A source of these rays had been observed in the eighteenth century by William Watson, an English apothecary and physician, when he passed electricity through a glass tube three feet long that had been partially emptied of air. Enchanted by the greenish yellow light that glowed within the length of the tube, he recorded," It was a most delightful spectacle to see the electricity in its passage." In 1875 William Crookes used a vacuum pump to remove almost all the air from a glass tube and discharged a high-voltage electrical current through it. Experimenting with the strange rays of light that came from the cathode, or negative plate of his apparatus, he found that he was able to bend the path of the glowing rays by placing a strong electromagnet near the tube. This simple experiment contradicted the most basic scientific principles. Light, as we have discussed in Chapter 2, is a form of radiation. It is energy, not matter. But magnetism is a property of matter, not energy. How could he reconcile the contradictory properties of these enigmatic "cathode rays"? He was never able to solve the maddening dilemma of the mysterious cathode rays. His fascinating apparatus, called the Crookes' tube or the **cathode ray tube**, attracted the best minds of the day to the scientific explorations that would lead to modern atomic theory. His discoveries altered our world in another significant way: the televisions and "CRT" screens of our computers are modern-day versions of the cathode-ray tube (Fig. 3-1).

The Cathode Ray Tube

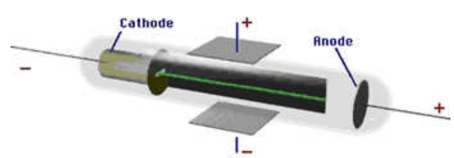


Fig. 3-1a. *A cathode ray tube. The stream of electrons from the cathode (negatively charged electrode) to the anode (positively charged electrode) interacts with gas molecules in the tube, creating a glowing beam.*

The Cathode Ray Tube

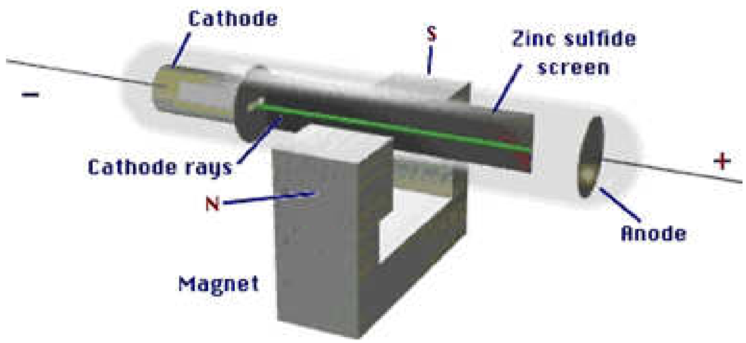


Fig. 3-1b. *When a magnet is brought near the cathode ray tube, the interaction of the electrons with the magnetic field deflects their path. If a fluorescing substance is placed in the path of the electrons, it glows. Glass fluoresces, so the glass tube will show a glowing spot where the electron beam hits.*

Fig. 3-1c. *Our TV picture tubes and computer "CRT" screens are also cathode ray tubes. An electron gun creates a spot of light, moving at great speed along parallel lines in the tube, giving the illusion of a continuous picture on the screen. Phosphorescing substances in the screen (called phosphors) are activated by the electrons and glow. Color television pictures are created by the combination of three electron guns, carrying the red, blue, and green components of the image. These activate dot-sized phosphors on the screen which give off red, blue, and green light.*

In 1895 the German physicist Wilhelm Konrad Roentgen, conducting experiments with a cathode ray tube covered with a shield of black cardboard, noticed an amazing phenomenon. A nearby sheet of paper painted with barium platinocyanide had begun to fluoresce. Invisible rays emanating from the cathode ray tube had penetrated both the glass walls of the tube and the opaque cardboard covering, imparting to the barium platinocyande the energy to produce the glowing emission of light that we call fluorescence. Roentgen soon found that these invisible rays, unlike the cathode rays, could easily penetrate through walls and even the human body. These mysterious rays, called **x-rays** or Roentgen rays, were recognized almost immediately to be a useful medical diagnostic tool.

Fig. 3-2. *Immediately after their discovery, x-rays became popular both for medical use and for parlor tricks. This is an illustration from a German postcard of an "x-ray party at the beach," exploiting the humor in the fact that Roentgen's x-rays could penetrate the modest beach attire of the day, exposing the skeletons of the partygoers. From Trevor I. Williams' "Science: A History of Discovery in the Twentieth Century."*

Immediately after their discovery x-rays became popular both for medical use and for parlor tricks. This is an illustration froma German postcard of an "x-ray party at the beach." From Trevor I. Williams' "Science: A History of Discovery in the Twentieth Century."

The French physicist Henri Antoine Becquerel was at this time studying the behavior of fluorescent and phosphorescent substances, which absorb light and then emit light of a longer wavelength. Might these substances also be emitting x-rays? While experimenting with uranium compounds and photographic plates, Becquerel discovered to his surprise that uranium compounds did indeed develop photographic plates, but not because of fluorescence, because they did not need to be exposed to light first. A uranium compound wrapped in black paper and accidentally left on a photographic plate left a spot on the plate. Even when uranium had been in the dark for months, a spot invariably appeared when it was placed on a photographic plate. Moreover, the rays emanating from the uranium compounds were not x-rays, because, unlike x-rays, they were electrically charged. What was the energy source of these rays? And might other elements besides uranium emit them?

Curie and Radioactive Elements

The brilliant young chemist Marie Curie recognized the potential significance of this phenomenon. Working in a ramshackle, unheated laboratory in Paris, she took on the ambitious task of examining all known substances for this phenomenon, which she called **radioactivity**. Early in her investigations she learned that the element thorium was also radioactive. Pitchblende ore, however, was so highly radioactive that neither radium not thorium could account for its activity. Was it possible that there was an element, as yet undiscovered, which was so highly radioactive that it could exhibit measurable radioactivity while in such low concentration that its existence had not been detected? Marie Curie felt certain that this must be so. Her husband Pierre abandoned his own prestigious experimental work to join her, and together they were soon able to announce that there were not one, but two new radioactive elements: polonium, named for Poland, Marie's native country, and radium, the most radioactive element of all. They undertook the back-breaking work of isolating these new elements from literally tons of pitchblende ore. After four years of exhausting labor in the unheated shed that served as their laboratory, they were successful. The Curies hold a unique place in the annals of science with their story of inspiring achievement and mutual devotion. One of the most touching images in scientific discovery is the description, by their daughter Irene, of Marie and Pierre Curie standing together in their dark laboratory late at night, enjoying the ghostly glow of the vials of radium. The Curies were jointly awarded the Nobel Prize with Becquerel. Marie continued her work after Pierre's tragic accidental death and was awarded a second Nobel Prize for her subsequent work with radioactivity.

Fig. 3-3. *Marie Curie was awarded two Nobel prizes for her work with radioactive elements.*

Thomson and the Electron

In England another brilliant young scientist was experimenting with mysterious rays. J.J. Thomson was determined to understand the cathode rays that had been observed by Crookes. After years of experimentation with his students, applying electrical and magnetic fields to the stream of cathode rays, Thomson changed atomic theory forever with his conclusion that the rays were composed of tiny negatively charged particles even smaller than atoms. These particles, which became known as **electrons**, had been present as a part of those few gas molecules which were left in the tube, and had been pulled away from their atoms. In other words, atoms were not the smallest possible units of matter, after all! From his experiments with electrical and magnetic fields, Thomson was able to determine the ratio of the charge to the mass of the electron, or the *e/m* where *e* is the charge on an electron and *m* is its mass.

By an ingenious yet simple experiment the charge of an electron was determined by Robert

Millikan at the University of Chicago. With an ordinary spray atomizer he produced a fine spray of oil drops in the area between two brass plates about one third of an inch apart which had been connected to the two poles of a powerful battery.. Holding a small tube of radium so that its rays would strike the oil drops and knock off electrons, he observed the effect on the oil drops with the aid of a microscope and a beam of light. When a drop lost an electron, it no longer fell like the other drops because it was attracted to the oppositely charged plate. Some droplets exhibited twice or three times the effect as they lost two or three electrons. "It was easy to see," Millikan wrote," that the slowest speed was the result of the loss of one electron. This proved conclusively that the smallest invisible load which I was able to remove from the droplet was exactly one electron and that all electrons consist of exactly the same quantity of negative electricity." Now that Millikan had determined the charge on an electron, it was possible to use Thomson's value of e/m to find the mass of an electron. It was eighteen hundred and fifty times less than the mass of a single atom of hydrogen.

Fig. 3-4*. Millikan's oil drop experiment enabled him to find the charge on an electron.*

The concept of the indivisible atom had now been shattered. Clearly there were yet smaller parts which composed the atom. Where inside the atom were these negatively charged electrons coming from? Since the atom itself was electrically neutral, somewhere in the atom there must be positively charged particles. What were these particles, and where were they inside the atom?

Rutherford and the Nucleus

The next major step in the discovery of the structure of the atom came in the laboratory of Ernest Rutherford in England. He was investigating the nature of the rays which emanated from radium, which he called alpha particles. Alpha particles, he found, were positively charged helium atoms. Their mass was eight thousand times as much as the mass of an electron, and their velocity as they were ejected from radium was an incredible twelve thousand miles per second. The alpha particle, he realized, was an ideal probe for exploring the atom. The breakthrough experiment, however, came in a very unexpected way. The suggestion came from Hans Geiger, a young German scientist working in Rutherford's laboratory who had discovered a way of observing and counting alpha particle radiation. As Rutherford recalled in later years,

> *... one day Geiger came to me and said "Don't you think that young Marsden (only twenty years old and still with no degree) whom I am training in radioactivity methods ought to do a small piece of research? Why not let him see if any alpha particles can be reflected from a solid surface and perhaps scattered through a large angle?*
>
> *Now I may tell you in confidence that I did not believe they would be, since we knew that the alpha particle was a very fast massive particle with a great deal of energy, and you could show that if the scattering was due to the accumulated effect of a number of small scatterings the chance of an alpha particle's being scattered backwards was very small. Then*

I remember Geiger two or three days later coming to me in great excitement saying, "We have been able to get some of the alpha particles coming backwards." It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you.

When aiming alpha particles at a piece of gold foil, chosen because gold can be beaten so thinly that it is only about a thousand atoms thick, Marsden had found that almost all the alpha particles passed through the foil or were deflected only very slightly. But a few of the alpha particles bounced right back at the observer. This could have happened only if they had hit something much more massive than themselves. Rutherford, at first amazed by these results, soon recognized their implications. Most of the mass of the gold atoms must be concentrated in one spot, the center of the atom. Since the negatively charged electrons were so light, this heavy center must be the place where the positive charge resides. Rutherford called this small but heavy positively charged center of the atom the **nucleus** when he formally announced his theory in 1911. According to Rutherford's calculations, the space occupied by the atom is about a million times greater than the space occupied by the nucleus.

Fig. 3-5. *The Rutherford experiment showed the existence of the nucleus. Most of the alpha particles pass through the gold foil; those that come near the nucleus of a gold atom are deflected.*

The basic features of the atom discovered through the efforts of Thomson, Millikan, Rutherford, and their colleagues remain in our mental picture of the atom today. At the center, infinitesimally tiny but unimaginable heavy, is the positively charged nucleus. Somewhere in the atom outside the nucleus in a volume a million times larger than the nucleus are the tiny, negatively charged electrons. The rest of the volume outside the nucleus in those atoms which comprise all the mass of the universe including those of the human body is made of ... empty space. Even today this concept, proven rigorously by Thomson's experiment, is difficult for us to accept philosophically. Surprisingly, though he believed in the indivisible atom, Lucretius poetically expressed the importance of the void in nature in *De Rerum Natura (0n the Nature of Things)*:

And yet all things are not held closely packed,
By body's nature hemmed on every side;

For there is void in things. This to have learned
Will bring thee rich reward...

Chadwick and the Nucleus

This mental picture of the atom is enough to produce considerable food for thought, especially among those who are philosophically inclined. It is not, however, quite complete. Rutherford and his fellow investigators were able to learn more about the composition of the nucleus. Bombarding nuclei with alpha particles, he was able to dislodge positively charged particles with the weight of a hydrogen atom, or almost two thousand times as heavy as the electron. Announcing his discovery in 1920, he named the new particle the **proton**. In 1932, bombarding beryllium with alpha particles, James Chadwick was able to knock out of the nucleus a particle with the same mass as the proton, but with no electrical charge. This new particle he called the **neutron**.

A Simple Model of the Atom

To summarize, then, the results of these experiments, a few decades into the twentieth century scientists knew that the atom is made of three types of particles: protons, neutrons, and electrons (Table 3-1). Protons and neutrons are found in the nucleus at the center of the atom, and they each weigh 1 atomic mass unit. The proton is positively charged, while the neutron is electrically neutral. Electrons are much lighter than protons and neutrons, and are found outside the nucleus. The electron has a negative charge. Atoms are electrically neutral because the number of positively charged protons is equal to the number of negatively charged electrons.

Table 3-1. Atomic Particles

Atomic particle	Mass (amu)	Charge	Position
Proton	1	+1	In the nucleus
Neutron	1	0	In the nucleus
Electron	1/1850	-1	Outside the nucleus

Atomic Number and Atomic Mass: Insights from Inside the Atom

The discovery of the proton, the neutron, and the electron as the basic units from which atoms are built was a great step forward in our understanding of the atom. To apply this knowledge, let us return to the list of the first twenty elements from Chapter 2. The order of the elements need no longer seem mysterious.

The atomic number of each element is the number of protons in the nucleus of the atom. For example, the atomic number of hydrogen is 1, corresponding with the fact that hydrogen has one proton in its nucleus. Similarly, helium, atomic number 2, has two protons, lithium, atomic number 3, three protons. **The atomic number gives not only the number of protons, but also the number of electrons in the neutral atom**, since the number of negatively charged particles, or electrons, must be exactly the same as the number of positively charged particles, or protons, in order for their electrical charges to balance.

The gram atomic mass of the elements seems a bit more complex to explain. Looking at the first twenty elements, we noticed that the gram atomic masses of the elements are seldom whole numbers, as the atomic numbers are. Moreover, the list did not always ascend to higher numbers in the same way that the atomic numbers did; Argon has atomic number of 18 and gram atomic mass 39.9, but potassium has atomic number 19 and gram atomic mass 39.1. Before the twentieth century, scientists were baffled by these seeming anomalies associated with gram atomic masses. Using our knowledge about the components of the atom, gram atomic masses of the elements can be explained by two simple facts. First, **the atomic mass of an atom is the number of protons plus the number of neutrons.** Each proton weighs 1 atomic mass unit, and each neutron weighs one atomic mass unit. So adding up the numbers of neutrons and protons gives the total mass of the atom in units of amu, or atomic mass units. The electrons are so much lighter than protons and neutrons (about two thousand times lighter) that their mass is insignificant in computing atomic masses.

Problem example 3-1: How many neutrons are present in an atom of oxygen, which has an atomic number of 8 and an atomic mass of 16 amu?

Since oxygen has an atomic number of 8, there must be 8 protons in an atom of oxygen. The weight of 8 protons is 8 amu. The weight of the neutrons in this atom must be 16 - 8 = 8 amu. Each neutron weighs 1 amu, so there must be 8 neutrons in an oxygen atom.

Fig. 3-6. *The nucleus of the oxygen atom has 8 protons and 8 neutrons, for a total mass of 16 amu.*

The second important fact we need to know in order to explain gram atomic masses is that elements can exist as different isotopes: **isotopes are atoms with the same number of protons, but different numbers of neutrons.** Because they have the same number of protons, isotopes must be the same element; the atomic number, or number of protons, defines which element an atom is. Because they have different numbers of neutrons, isotopes have different masses.

If more than one isotope of an element is present, as is normally the case, the gram atomic mass of the element is the weighted average, expressed in grams, of the isotopes that are present. Notice that *the mass of an atom is expressed in units of atomic mass units, or amu*. Individual atoms are much too small, of course, to weigh out. Since grams are the units used to measure weights in the laboratory, the relative masses of the atoms are expressed as grams when calculating how much of one substance we need to react with another. In amu or in grams, to get equal numbers of hydrogen atoms and oxygen atoms, the mass of the oxygen must be sixteen times as much as the hydrogen, because the oxygen atom is sixteen times heavier than the hydrogen atom. The atomic masses we find listed on the periodic table are usually given as *gram atomic masses, or atomic masses expressed in grams*.

Fig. 3-7*. The gram atomic mass of an element is the weighted average of the naturally occurring isotopes of that element. For example, the gram atomic mass of chlorine is approximately 35.5 because about 75% of chlorine atoms have a mass of 35 amu (17 protons, 18 neutrons), and about 25% of chlorine atoms have a mass of 37 amu (17 protons, 20 neutrons).* Drawing: 100 chlorine atoms (spheres): 75 labelled 35 amu, 25 labelled 37 amu; if possible, some representation of protons and neutrons, in which case the spheres represent nuclei.

Problem example 3-2: The gram atomic mass of carbon is 12.011 g. What is the mass in amu of the atom of the most common isotope? How many neutrons in this isotope? What can you say about the other existing isotopes of this element?

The nearest whole number to the gram atomic mass is 12, so the mass of the atoms of the most common isotope must be 12 amu. The atomic number of carbon is 6, meaning there are 6 protons in this atom. The number of neutrons is 12 - 6 = 6. Since the gram atomic mass is slightly heavier than 12, the weight of the most common isotope, there must be another isotope occurring less often, which is heavier than the most common isotope; one possibility is an isotope with 7 neutrons instead of 6.

Fig. 3-8. *Most carbon atoms have 6 protons and 6 neutrons in the nucleus, for a total mass of 12 amu.*

Now we can finally explain the anomaly we noticed in the list of the first twenty elements in Chapter 2. Before we knew about the existence of protons, neutrons, and electrons, the major piece of quantitative information we had about the elements was their gram atomic masses. The first twenty elements were listed in order of their increasing masses, until we came to argon and potassium. Argon is listed as element 18, with gram atomic mass of 39.948, and potassium as element 19, with the smaller gram atomic mass of 39.098. Attempting to organize the elements on the basis of weight alone was the source of great frustration for scientists for centuries. Understanding the properties of the neutron explains seeming discontinuities like the masses of argon and potassium.

Problem example 3-3: Find the number of protons and the number of neutrons in the most common isotope for the elements argon and potassium.

Argon, with its atomic number of 18, has eighteen protons. The nearest whole number to its gram atomic mass of 39.948 is 40. The most common isotope of argon must have 40 - 18 = 22 neutrons. Potassium, with its atomic number of 19, has nineteen protons. The nearest whole number to its atomic mass of 39.098 is 39. The most common isotope of potassium must have 39 -19 = 20 neutrons.

Since the position of an element on the periodic table is determined by the number of protons, not the mass (protons plus neutrons), it is possible as with argon and potassium to have atomic masses slightly "out of order."

Bohr's Broken Rainbow and the Electron's Position

The nucleus determines the mass of the element, but the protons and neutrons of the nucleus are locked deep in the center of the atom. The electrons, located outside the nucleus, are the atomic particles most likely to interact with other atoms. It is this interaction of electrons from different atoms which can lead to the combination of atoms with one another to form molecules, or even the loss or gain of electrons to form positively or negatively charged particles called ions. When learning about chemical reactions, then, we will be learning about the interactions of electrons, those lightweight negatively charges particles outside the nucleus.

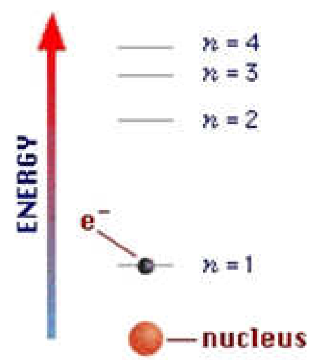
So far our mental picture of the electrons in the atom is vague. The neutrons and protons are inside the nucleus, an area both small and well-defined in its position. The position of the electrons has been described only as somewhere outside the nucleus. It is as if a college student were to be informed that her English class is in Building 8, room 218, and her French class is ...somewhere else within the college. How could anything be located with such a hopelessly vague address? The question of the location of the electrons was an especially puzzling problem, and the first scientist to provide the key to this puzzle was a young Danish scientist named Niels Bohr. In 1913, only one year after joining Rutherford's research group, he published his bold conception for the structure of the hydrogen atom. Applying Max Planck's revolutionary theory of quantized energy, he developed a theory of electron energy levels that explained the way in which hydrogen emits light.

Fig. 3-9. *The bright colors of fireworks display light being emitted from the electrons of atoms. Bohr's theory helped to explain why different elements give different colors.*

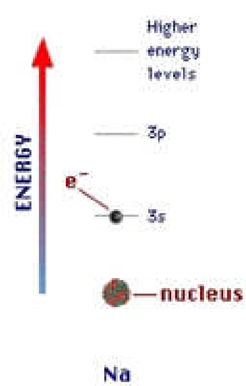
The emission of light from atoms is easy to observe in the laboratory. Whenever salts are placed in a flame, bright colors are observed that are characteristic of the elements: red for strontium, orange for calcium,and yellow for sodium, for example. At home, whenever cooking water containing salt boils over on a gas stove, the characteristic yellow color from sodium appears in the flame. Brightly colored fireworks depend on the presence of certain elements to produce flashes of light of the desired color.

Fig. 3-10. The light emitted by an excited hydrogen atom appears as a series of lines at characteristic wavelengths. The light emitted by an excited sodium atom appears as another set of lines, corresponding to wavelengths of light.

Electron energy levels in hydrogen



Fireworks: energy transitions in sodium



When viewed through an apparatus called a spectrograph, the colored light emitted by an element is spread out and reveals not a continuous rainbow-type spectrum, but a pattern of very distinct, narrow lines appearing only at predictable characteristic wavelengths. Why does each element emit light only at a few wavelengths instead of across the whole rainbow of possible colors? Bohr proposed an answer: Electrons can exist only at certain energy levels within the atom. In the language of physics the energy of the electron is quantized, or discontinuous. In order to go to a higher energy, an electron must gain exactly enough energy to jump from its normal, or **ground state**, level to a higher, or **excited** energy state. It is not possible for the electron to stop somewhere in between. When an electron loses energy, it must lose exactly the difference in energy between its excited state and the ground state energy level as it jumps to the lower level. This emitted energy corresponds to a specific wavelength of light, which often falls into the visible range (Chapter 2). The line spectra emitted by the elements are the result, then, of the energy differences between the electron energy levels in their atoms.

The Bohr model and the Quantum model of the atom

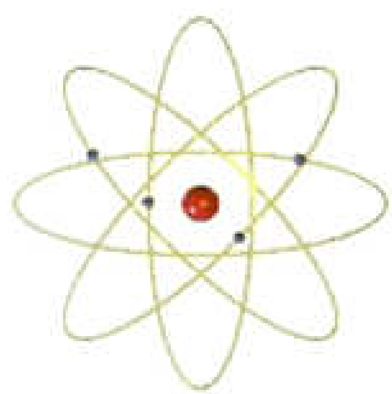


Fig. 3-11. *Bohr's model of the atom pictured the electrons as being positioned in energy levels like concentric spheres. In cross section, the energy levels look like circular "tracks." When energy is absorbed by an electron, it jumps up to a higher energy level. When it falls back down to a lower energy level, energy is emitted.*

The energy levels in the Bohr model of the atom can be envisioned as concentric shells, or circular tracks as seen in cross-section. The electrons travel around the central nucleus in these orbitals much as the planets travel around the sun in our solar system. Each energy level can contain only a limited number of electrons. The maximum number of electrons that can occupy an energy level is given by the formula $2n^2$, where n is the number of the energy level. For example, the first energy level, the one closest to the nucleus, can contain $2(1)^2$, or two electrons.

Problem example 4: What is the maximum number of electrons in energy level 2?

$2(2)^2 = 2(4) = 8$ electrons

Problem example 5: How many electrons are to be found in each energy level of an atom of chlorine?

To find the number of electrons in each energy level, begin at the lowest level and fill each shell in turn, keeping in mind the maximum number each shell can hold.
Chlorine, atomic number 17, has 17 electrons:

2 electrons in level 1
8 electrons in level 2
7 electrons in level 3

The Quantum Atom: A New Concept

Although it has been superseded by a more sophisticated theory, the Bohr atom remains to this day a useful model for visualizing electrons within the atom. The limitations of this theory, however, soon became evident. Bohr's equations predicted correctly the emission spectrum for the hydrogen atom, but could not be applied successfully to other elements. In order for a true picture of the world of the atom to be developed, Bohr's model of the atom had to be replaced by an even more revolutionary one.

Bohr's simple model of electron energy levels explained the characteristic colors of the emission spectra of elements as electrons jumped from one energy level to another, consistently emitting light that corresponded to the energy difference between the two energy levels. To create this model, however, he first had to postulate that the world inside the atom had rules all its own; the well-known rules of physics simply did not apply in this domain. For example, the nucleus has a positive electrical charge and the electrons have a negative charge; since opposite charges attract, why didn't the electrons fall right into the nucleus instead of staying in their energy levels? Electrons could exist only in specified energy levels, and never in-between; it was as if they lived in a building and always had to take the elevator, never the stairs; but why? A new field of physics called quantum physics developed in response to these questions.

Classical physics taught that matter and energy were totally different phenomena. **Quantum mechanics,** the new theory of the behavior of electrons within the atom, shows that electrons within the atom have the properties of both waves and particles. Though difficult to visualize, this theory does have a certain logic. As in Chapter 2, only certain wavelengths can fit into a confined space (Fig. 2-9), or form standing waves with a rope. Hence, if an electron exhibits wave properties, when confined within the space of an atom, it can only have certain wavelengths and their corresponding energies. Erwin Schrodinger in 1926 succeeded in expressing the wave behavior of electrons within the atom mathematically in the **Schrodinger equation**. His quantum mechanical equations showed that energy levels were not the only quantized properties of the atom, and that *n*, the number of the energy level, was only one of three quantum numbers needed to specify an electron's place in the atom. Each electron is specified by a unique combination of these three quantum numbers.

In the quantum mechanical model of the atom, the position of the electron is described as a probability distribution, or orbital, which describes the probability of finding the electron in a given place in the atom. Quantum number *l* determines the shape of the orbital, which can be described as *s, p, d, or f. s* orbitals are spherical, reminiscent of the Bohr atom model, but energy levels are not all spherical; each orbital type has a different characteristic shape. *p* orbitals, for example, are dumbbell-shaped.

Fig. 3-12. The quantum mechanical model describes the position of the electron in the atom in terms of the probability of finding it in a given area. The electron on a hydrogen atom, a 1s electron, is represented here in (a)an electron density distribution showing probabilities of finding the electron at different distances from the nucleus, and (b)in a contour representation, showing a sphere within which the electron can be found with a 90% probability.

Fig. 3-13. The p orbitals are dumbbell-shaped. There are three p orbitals, oriented at right angles to each other along the three coordinate axes x, y, and z.

The electron spins about a central axis like a top. This spin can be in one of two directions, described by the spin quantum number m_s. The possible values of m_s are +1/2 and -1/2.

Fig. 3-13. The electron spins like a top around a central axis. The two possible spin directions are described by the two possible values +1/2 and -1/2 of the spin quantum number m_s.

Table 3-2. The Three Quantum Numbers

Quantum Number	*Possible Values*	*Physical Significance*
n	1,2,3,4,5	Energy level
l	0 up to n-1	Orbital type
m_s	+1/2, -1/2	Direction of electron spin

The implications of quantum mechanical theory were many. The **Heisenberg uncertainty principle**, which states that it is impossible to determine both the energy and position of an electron at the same time, had profound philosophical implications. Always before, science had pursued certainty in defining nature. For the first time a law defined that there were limits to certainty.

The insights of quantum mechanics have played a major role in the science of the twentieth century. Quantum mechanical calculations, however, are known for their complexity. We will find that simpler insights: the basic components of the atom, and the existence of electron energy levels, will meet most of our needs in explaining the way atoms interact to form the substances of our world.

CONCEPTS TO UNDERSTAND FROM CHAPTER 3

The cathode ray tube, in which a high voltage is applied across an evacuated tube, was an early tool in exploring the atom. The "ray" observed in the tube was found by Thomson to be negatively charged electrons from the negatively charged electrode,or cathode, of the cathode ray tube.

The mass of an atom is concentrated in its nucleus.

Protons, with a charge of +1 and a mass of 1 amu, are found in the nucleus.

Neutrons, with a mass of 1 amu and no charge, are found in the nucleus.

Electrons, with a charge of -1 and a mass of 1/1850 amu, are found outside the nucleus.

The atomic number of an atom is equal to the number of protons in the nucleus.

The number of electrons is equal to the number of protons in the neutral atom.

The atomic mass of an atom in amu (atomic mass units) is equal to the number of protons + the number of neutrons in the nucleus.

Isotopes are atoms with the same number of protons but different numbers of neutrons.

The atomic mass of an element is the weighted average of the masses of the naturally occurring isotopes.

The electrons occupy energy levels within the atom. The number of electrons in each energy level is $2n^2$ where n is the number of the energy level.

Electrons absorb energy when going from a ground state (their normal energy level) to an excited state (higher energy level). They give off energy when falling from an excited state to a ground state.

Quantum mechanics, the theory of the behavior of electrons within the atom, states that electrons within the atom have the properties of both waves and particles.

Each electron is specified by a unique combination of the three quantum numbers n, l, and m_s. (See Table 3-2 for their possible values and physical significance.)

The **Heisenberg uncertainty principle** states that it is impossible to determine both the energy and position of an electron at the same time.

FACTS AND NAMES TO LEARN FROM CHAPTER 3

William Crookes developed the cathode ray tube, which was an early tool in exploring the atom's structure.

Wilhelm Konrad Roentgen discovered that penetrating x-rays were a by-product of the cathode ray tube.

Henri Becquerel discovered radiation emanating from uranium ore.

Marie Curie and her husband Pierre researched the phenomenon of radioactivity and discovered several radioactive elements.

J.J. Thomson discovered that the cathode ray was made of negatively charged electrons. He determined the charge-to-mass ratio for the electron.

Robert Millikan found the charge on an electron with a simple oil-drop experiment.

Ernest Rutherford discovered the nucleus by bombarding gold foil with alpha particles.

James Chadwick discovered the neutron.

Neils Bohr discovered the existence of electron energy levels in the atom.

Erwin Schrodinger derived a mathematical equation which describes the wave behavior of electrons in the atom. Quantized energy levels and sublevels as well as other properties of the electron are described in related equations.

SKILLS TO ACQUIRE FROM CHAPTER 3

After finishing this chapter, you should be able to:

Be able to predict the number of protons and electrons in an atom if given the atomic number.

Be able to predict the most common naturally occurring isotope of an element if given the atomic number and the gram atomic mass.

Be able to predict the maximum number of electrons in each energy level of an atom from the formula $2n_2$.

Be able to predict the number of electrons in each energy level for the first eighteen elements.

Name ___________________________________ **Date _____________**

PROBLEMS TO SOLVE USING CONCEPTS, FACTS, AND SKILLS FROM CHAPTER 3

3-1. For the element helium, with atomic number 2 and atomic mass 4.0, find the following:

a. The number of protons

b. The number of electrons

c. The number of neutrons in the most commonly occurring natural isotope.

d. The number of electrons in the first electron energy level.

3-2. For the element oxygen, with atomic number 8 and atomic mass 16.0, find the following:

a. The number of protons

b. The number of electrons

c. The number of neutrons in the most commonly occurring natural isotope.

d. The number of electrons in the first electron energy level.

e. The number of electrons in the second energy level.

3-3. For the element sulfur, with atomic number 16 and atomic mass 32.1, find the following:

a. The number of protons

b. The number of electrons

c. The number of neutrons in the most commonly occurring natural isotope.

d. The number of electrons in the first electron energy level.

e. The number of electrons in the second energy level.

f. The number of electrons in the third energy level.

3-4. For the element argon, with atomic number 18 and atomic mass 39.9, determine the following:

a. The number of protons

b. The number of electrons

c. The number of neutrons in the most commonly occurring natural isotope.

d. The number of electrons in the first electron energy level.

e. The number of electrons in the second energy level.

f. The number of electrons in the third energy level.

3-5. In your own words, describe the experiment in which the nucleus was discovered.

3-6. In 1926 a famous dinner was given in Cambridge, England in the honor of J.J. Thomson where a song was sung by the attendees. To what particle does this verse refer?

> Though Crookes at first suspected my presence on this earth
> Twas J.J. that found me -- in spite of my tiny girth.
> He measured first the "e by m" of my electric worth:
> I love J.J. in a filial way, for he it was gave me birth.

3-7. How is a television screen related to a Crookes tube?

3-8. An alpha particle has a mass of 4 amu and a charge of +2. Which element on the periodic table has a nucleus with the same properties?

3-9. The symbol of radium, the radioactive element studied by Marie Curie, is taken from the first two letters of its name. From the periodic table of the elements in Chapter 5, find:

a. The symbol of the element

b. The atomic number

c. The atomic mass

d. The number of protons in the nucleus

e. The number of neutrons in the most common naturally occurring isotope.

3-10. Calculate the maximum number of electrons in energy level 4.

3-11. Match each symbol with the correct descriptive phrase by placing the correct letter in the blank.

_____ n a. Describes electron spin

_____ l b. Describes orbital type of an electron

_____ m_s c. Describes energy level of an electron

_____ s d. A dumb-bell shaped orbital type

_____ p e. A spherical orbital type

3-12. Match the discoverer with the discovery by placing the correct letters in the blanks.

_____ Millikan	A. Radium; radioactivity studies
_____ Curies	B. Radioactivity of uranium
_____ Schrodinger	C. The neutron
_____ Crookes	D. The nucleus
_____ Becquerel	E. Cathode ray tube
_____ Roentgen	F. X-rays
_____ Bohr	G. Charge on an electron
_____ Rutherford	H. The electron; e/m
_____ Thomson	I. Electron energy levels
_____ Chadwick	J. Quantum theory

3-12. Fill in the blank the name of the correct element.

a. ______ has 4 electrons.

b. ______ has 4 neutrons in the most common naturally occurring isotope.

c. ______ has 4 protons.

d. ______ has 9 electrons.

e. ______ has a half-filled first electron energy level.

3-15. Compare the Bohr model and the quantum mechanical model of the atom in terms of their descriptions of the locations of the electrons within the atom.

3-14. Fill in the blanks in the table below.

Atomic particle	*Mass (amu)*	*Charge*	*Position*
___________	1	+1	
Neutron	_________	______	In the nucleus
___________	1/1850	-1	Outside the nucleus

RECOMMENDED READINGS, CHAPTER 3

"Crucibles, the Story of Chemistry from Ancient Alchemy to Nuclear Fission," Fourth Ed., Bernard Jaffe, Dover Publications, 1976.

"Madame Curie," by Eve Curie, translated by Vincent Sheehan, DaCapo, 1986.

"Science: A History of Discovery in the Twentieth Century," Trevor I. Williams, Oxford University Press, 1990.