Chapter 8: Models for Atoms

In the previous chapters we have had frequent occasion to contrast direct observations of nature with the models scientists have made for interrelating these observations. The many-interacting-particles (MIP) model for matter has been a powerful tool to this end. In this model, matter is composed of micro-domain particles that are in motion and that interact-at-a-distance with one another by means of an intermediary field. During the nineteenth century the various types of energy—kinetic, thermal, elastic, chemical, phase, electromagnetic field (including radiant), and gravitational field—were investigated intensively. As we pointed out in Section 4.5, and as we will describe in greater detail in later chapters, it is possible to explain thermal, elastic, chemical, and phase energies in terms of the kinetic energies of all the particles and the field energy arising from their interaction.

During the twentieth century, physicists have set themselves the goal of explaining all macro-domain phenomena, including perhaps even life, in terms of these particles, which are called atoms or molecules. Great strides in formulating these explanations have been made, except in the case of phenomena involving the gravitational field. Even Einstein, who reformulated the gravitational interaction in a very novel and general way in his general theory of relativity, was not able to relate gravitation effectively to the electric, magnetic and other fields.

One of the most notable areas of progress has been in the invention of models for atoms themselves. Instead of being conceived of as simple point-like particles with no internal structure, atoms are now viewed as complex systems composed of simpler constituents. The properties of atoms are explained in terms of the arrangement and motion of the constituents. In this chapter we will review some of the studies of the current century that have led to the presently accepted models for atoms.

8.1 The electrical nature of matter

Dalton's atomic theory. When John Dalton proposed his atomic theory of chemical reactions, the particles were called *atoms* (from the Greek word for "indivisible") because they were conceived of as being ultimate constituents that would permit no further subdivision. In this view, which gradually became accepted during the first half of the nineteenth century, each chemical element is composed of a different kind of atom. There were about 90 kinds of atoms, each with its own characteristics (Fig. 8.1). The atoms were believed to have properties that could account for the chemical activity and various other macro-domain properties (hardness, appearance, melting and boiling temperatures, and so on) of specimens of the element.

"... the existence of these ultimate particles of matter can scarcely be doubted, though they are probably much too small ever to be exhibited by microscopic improvement. I have chosen the word atom to signify these ultimate particles...."
John Dalton
A New System of Chemical Philosophy, 1808

Electric conduction in solid and liquid materials. As soon as the existence of Dalton's atoms became non-controversial, questions arose about the "intrinsic" properties of the atoms and whether models could be constructed to account for them. In other words, scientists asked in what way an oxygen atom differed from a nitrogen atom, why solid copper conducts an electric current while solid sulfur does not, and how

Humphry Davy (1778-1829) was director of an institution in Bristol, England, that investigated the medicinal properties of various gases when he published the work on nitrous oxide (1800) that made him famous. He was appointed to the Royal Institution in London, where he continued his research. Davy is remembered as the inventor of the coal miner's safety lamp, and he was the first to explain electrolysis and to suggest that electricity and chemical interactions are due to the same ultimate cause (the electric field, as we now call it)*

Michael Faraday (1791-1867), a blacksmith's son, was apprenticed in 1804 to learn the bookbinding trade. The boy used his spare time to read books on electricity and chemistry, and in 1812 he attended a series of lectures by Humphry Davy. Faraday wrote a summary of the lectures and sent it to Davy, along with a request for a job. The summary must have been brilliant; it attracted Davy's attention, and he appointed Faraday as a laboratory assistant at the Royal Institution. Faraday eventually succeeded Davy as Director, and Davy was to boast that of all his discoveries the greatest was Michael Faraday. Many historians rank Faraday as the greatest of all experimental physicists. Although he knew no mathematics, he invented the field concept and formulated a complete descriptive theory of electricity.

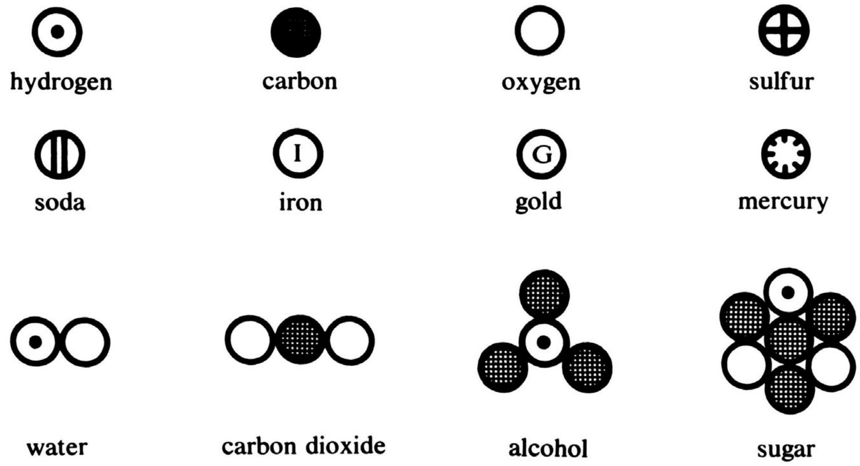


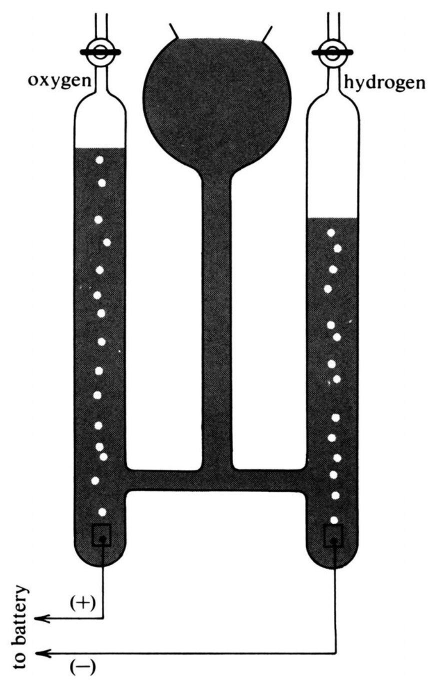
Figure 8.1 Dalton's symbols for the atoms of common elements. The bottom row shows his models for the more complex "atoms" of common compounds.

the metal sodium and the gas chlorine could interact to produce the white crystalline substance sodium chloride (ordinary table salt). Furthermore, a solution of sodium chloride and water conducts an electric current, but solid sodium chloride and pure water separately do not.

Electrolysis. Sir Humphry Davy and Michael Faraday found another effect: melted table salt (sodium chloride) can be decomposed into sodium and chlorine by the passage of an electric current, but the element copper is not modified on the macro level by an electric current. The decomposition of sodium chloride is an example of *electrolysis* (using electricity to divide something into its parts). An electric current can

Figure 8.2 (to the right) The electrolysis of water to form hydrogen and oxygen gases. Note the ratio of the volumes of gases produced. What do you conclude about the composition of water? On this basis, can you suggest a modification of Dalton's symbol for water above in Figure 8.1?

It is interesting that Dalton knew about this experiment, and others giving similar results about the simple whole number ratios of the volumes of combining gases, but he did not accept these findings, possibly because the particular model of atoms he had developed couldn't explain it!



OPERATIONAL DEFINITION The quantity of electric charge is measured by the mass of hydrogen it liberates in the electrolysis of water. The unit of electric charge, called the Faraday *(symbol F), is the quantity of electricity associated with 1 gram of hydrogen.*

Sir Joseph John Thomson (1856-1940) was born in Manchester and studied at Cambridge, England. In 1884 he was appointed to what was then the most prestigious position in physics, the Cavendish chair of physics at Cambridge. His discovery of the electron and its properties in 1897 won him a Nobel Prize in 1906.

"I hope I may be allowed to record some theoretical speculations ... I put them forward only as working hypotheses, ... to be retained as long as they are of assistance ... The phenomena in these exhausted tubes reveal to physical science a new world—a world where matter may exist in a fourth state, where the corpuscular theory of light may be true, and where light does not always move in straight lines, but where we can never enter, and with which we must be content to observe and experiment from the outside."
Sir William Crookes
Philosophical Transact., 1879

also break up (electrolyse) water, producing hydrogen gas and oxygen gas in a volume ratio of two to one (Fig. 8.2). Faraday concluded from these observations that atoms of matter must be endowed with electrical charges. In fact, the mass of material produced in electrolysis can be used to formulate an operational definition of the quantity of electric charge; the details of the definition are in the left margin.

Franklin's electric fluid. Even air permits the passage of an electric current, as in lightning or an electric spark. Certain materials can be given an electric charge by rubbing. Benjamin Franklin studied these phenomena and concluded, as explained in Section 3.5, that there was one electric fluid whose presence in greater or lesser amounts showed up as positive and negative charges. According to Faraday, these electric charges had to originate within the atom. The generation of electromagnetic waves in association with sparks (Section 7.3) was further evidence that the constituents of matter had electrical properties. Arrayed against these conclusions was the lack of electric effects in ordinary pieces of matter such as a glass of water, a coin, or the air we breathe.

Electric conduction in air. William Crookes (1832-1919) and J. J. Thomson studied in detail the conduction of electric current through air in a glass tube. Fig. 8.3 shows their apparatus. Electric terminals were connected to metal pieces (called electrodes), which were sealed through the ends of the tube. The gas in the tube could be pumped out so as to produce a partial vacuum inside. At ordinary atmospheric pressure, sparks jumped from one electrode to the other. At low pressure, the gas became luminous. But at very low pressure, the glass of the tube itself glowed—evidence of interaction-at-a-distance between the electrodes and the glass. When a metal screen with a fluorescent covering was placed in the tube, it glowed and cast a "shadow" on the glass tube, presumably by intercepting the "rays" that were passing between the electrodes and the glass. In this way it was possible to show that the negative electrode, called the *cathode,* was the source of the rays, which were called cathode rays, and the apparatus became known as a "cathode ray tube." Whether the rays were a stream of particles or a wave phenomenon was not yet known.

Figure 8.3 (to right) An electric current can pass through low-pressure air in a glass tube, giving rise to many interesting luminous effects. This simple "cathode ray tube" stimulated many fruitful investigations, giving birth to the electron, electronics, the TV tube, X-rays, and other discoveries.

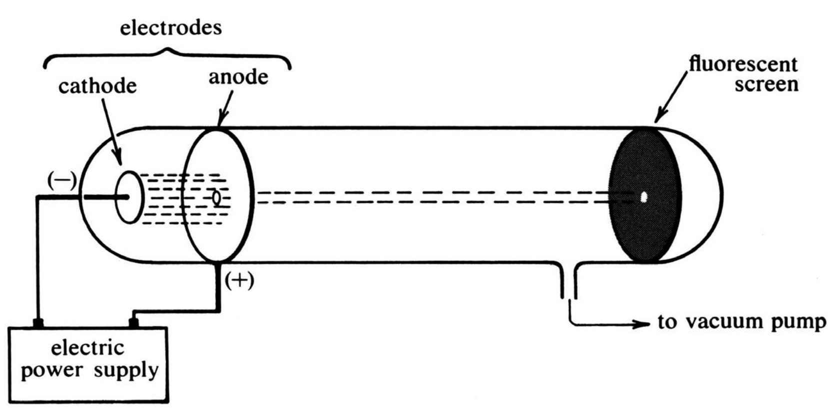
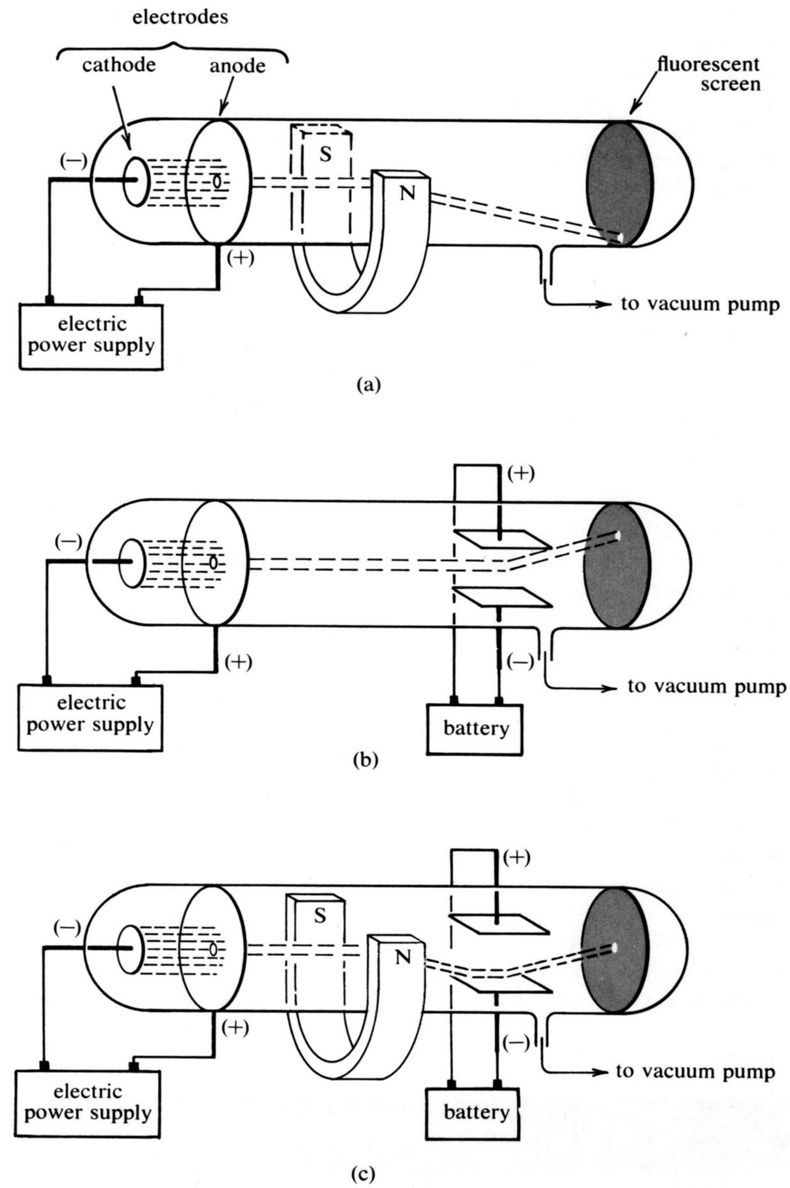


Figure 8.4 (to right). Deflection of cathode rays by magnetic and electric fields. (a) Deflection by a magnetic field. (b) Deflection by an electric field. (c) Deflection by both magnetic and electric fields.

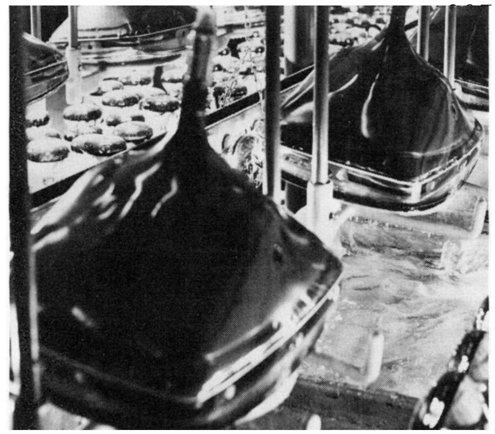


"... The most diverse opinions are held as to [cathode] rays; according to the almost unanimous opinion of German physicists they are due to some [wave] process in the aether to which—inasmuch as in a uniform magnetic field their course is circular and not rectilinear—no phenomenon hitherto observed is analogous; another view of these rays is that, so far from being wholly aetherial [wave-like], they are in fact wholly material, and that they mark the paths of particles of matter charged with negative electricity . . . I can see no escape from the conclusion that they are charges of negative electricity carried by particles of matter. The question next arises, What are these particles? are they atoms, or molecules, or matter in a still finer state of subdivision?"

J. J. Thomson
Philosophical Magazine, 1897

Electrons. By collecting the rays, letting them interact with electrically charged bodies, deflecting them with magnets (Fig. 8.4), and quantitatively measuring the effects, J. J. Thomson was able to show that the cathode rays carried electric charge and inertial mass in a fixed ratio. He therefore proposed a particle model to explain the interaction-at-a-distance between the cathode and glass: many tiny, identical micro-domain particles with a definite mass and negative charge are emitted by the cathode and acquire kinetic energy on being attracted by the positive electrode (the *anode).* Thomson adopted the term *electron* for these particles. Robert Millikan later measured the electric charge of a single electron in 1909 by means of his ingenious oil drop experiment, the first measurement of a physical

Figure 8.5 Color television tubes in an assembly line are being baked over heat lamps to bond a special phosphorescent material to the inside of the glass screen. TV tubes are essentially cathode ray tubes; the electrons are ejected by the cathode (at top of photograph), formed into a beam, and allowed to strike the screen (at bottom), transferring some of their energy to the phosphorescent material, which emits light of specific colors. Magnetic fields are used to direct the electron beam to specific points on the screen and thus form the image we see.



quantity in the micro domain. Millikan's value of the electric charge plus Thomson's value for the charge-to-mass ratio yielded the actual mass of the electron. Modern application of the cathode ray tube has created the electronics industry and revolutionized communications (Fig. 8.5).

Robert A. Millikan (1868-1953), a physicist at the University of Chicago and an experimentalist of extraordinary talent and patience, developed his "oil drop experiment" by which changes of one electron charge could be detected. He was the first to carry out a direct measurement of the charge of the electron. He also verified Einstein's hypothesis of light quanta. His painstaking work on the charge of the electron and the photoelectric effect won him the Nobel Prize in 1923.

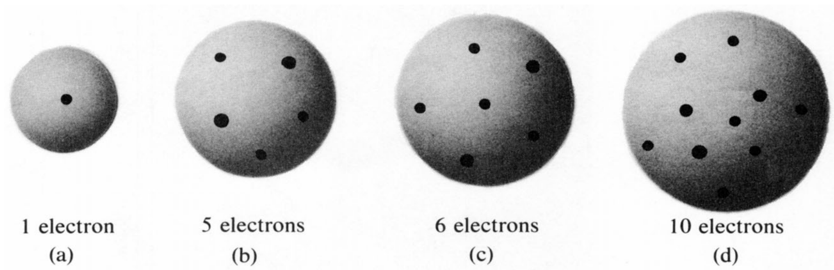
Millikan found that the charge of one electron was 1.7×10^{-24} *faraday. Hence 1 gram of hydrogen has* 6.0×10^{23} *electrons associated with it.*

Electron waves. According to J. J. Thomson in 1897, cathode rays were minute, electrically-charged "corpuscles." However, physicists in Germany, where Crookes had originally discovered cathode rays, thought of the cathode rays as similar to light and other types of electromagnetic waves (including the recently discovered X-rays), and they attempted to explain the cathode rays on the basis of a wave motion of the "aether" (see quote from Thomson on previous page). Thomson overcame these arguments by the turn of the century with a masterful series of experiments, volumes of data and many outstanding papers demonstrating how his data confirmed the predictions of the particle model and conflicted with those of the wave model.

As a matter of fact, no one at the turn of the century actually tested the cathode rays, directly, for wave properties. One reason for this was Thomson's experiments and tightly-woven arguments; another was the fact that electric charges had always been associated with matter, and no one had ever observed charges transported from place to place by waves. In any case, no one pursued this for almost 30 years, until, as we explain below in Section 8.4, a wave model for all matter was proposed on theoretical grounds. As a result, in 1926, Davisson and Germer performed an experiment in which cathode rays were intercepted by a nickel crystal, which had regularly arranged atoms that functioned like a diffraction grating. Davisson and Germer found a diffraction pattern, as predicted by a wave model for the cathode rays! However, the wavelength of the electrons was extremely short, much shorter than the wavelength of visible light.

The present view, therefore, is that the propagation of electrons is best described as a wave phenomenon. In most situations, the wavelength is very short, however, so that diffraction and other wave-like behavior appear only with micro-domain-sized slits, as in the nickel crystal. Thus Thomson's experiments, with macro-domain size slits, did not reveal wave effects. The modern theory includes the possibility of forming micro-domain wave packets from electron waves. These tiny wave packets, which carry mass and electric charge, behave like particles in macro-domain experiments.

Figure 8.6 (to right). J. J. Thomson's "plum pudding" model for the atom. The tiny electrons (the raisins) repel one another but are held in place (in mechanical equilibrium) by a much larger "pudding" of positive charge. The positive charge has most of the mass of the atom but, unlike the electrons, is like a diffuse "cloud" spread throughout the volume of the atom. The electrons, if stimulated, can vibrate in place to produce light waves.



The MIP model. At the beginning of the twentieth century it was definitely established by experiment (electrolysis, the ability of matter to emit electromagnetic waves, and cathode ray studies) that matter could be divided into electrically charged components. Presumably, therefore, each electrically neutral atom had positive and negative parts. At last it was possible to identify how the particles interacted in the MIP model: the interactions were electric, transmitted by means of an electric field. Macro-domain chemical, phase and elastic energy were, therefore, really just electric field energy of the atoms and molecules. The mutual attraction of the positive and negative parts in adjacent atoms could explain the structure of molecules, crystals, and liquids. The relative motion of the electrically charged parts could explain the emission of electromagnetic waves (Section 7.3), including the spectra of visible light.

The negative constituents of atoms were the very light electrons, which Thomson had measured as less than one thousandth the mass of the atom as a whole. What was the positive part, which appeared to have almost all the mass of the atom?

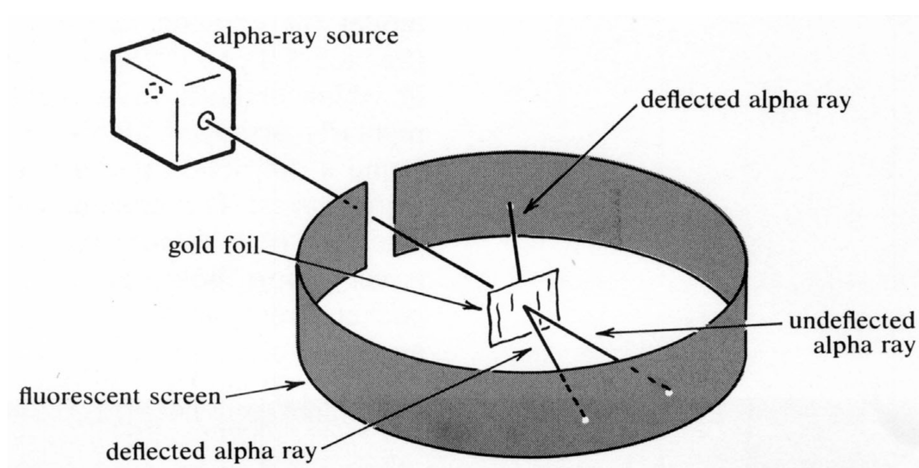
8.2 Early models for atoms

Thomson's model. J. J. Thomson himself was one of the first to propose an atomic model that took into account the atom's electrical nature (Figure 8.6). He suggested that an atom consists of a positive electric charge that is uniformly distributed within a spherical region of about

Figure 8.7 (below). Ernest Rutherford's scattering experiment. The alpha particles, which were known to have much more mass than an electron, emerge from the source in a narrow beam at high speed and strike the gold foil target. Most of the alphas continue undeflected through the foil in a straight line, but some are deflected at various angles. The physicist sits in darkness for long periods, counting the bright flashes on the circular screen created by the impacts of the deflected alpha particles.

Rutherford found that the number of alphas he counted at various angles did not match the predictions of the Thomson model (above). In particular, Rutherford was astounded (see quote on next page) to find a very few flashes showing that the alphas were occasionally bouncing <u>backwards</u>. He initially thought that such flashes could not be genuine.

It is extraordinarily unlikely that Thomson's plum pudding model (above) would bounce a heavy alpha particle backwards. Thus Rutherford invented another model, with the positive charge concentrated in a tiny but heavy <u>nucleus.</u>



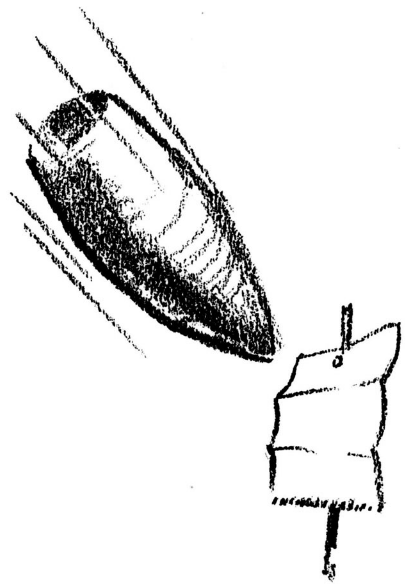


Figure 8.8 Analogue model for the collision of an alpha ray with an electron: a 15-inch shell hits tissue paper, as suggested by Ernest Rutherford in quote below to left. Would you expect the shell (or the alpha ray) to bounce back? Estimate the ratio of masses - this gives a rough estimate of the probability that the shell would bounce back.

Ernest Rutherford (1871-1937) was a New Zealander who came to England on a scholarship in 1895. After three years as a student of J. J. Thomson at Cambridge, he accepted a professorship at McGill University, Canada. Rutherford showed such virtuosity as an experimentalist that he was brought to Manchester as director of a research laboratory in 1907. At Manchester, Rutherford developed his nuclear model of the atom. An extraordinarily warm, perceptive, and stimulating man, Rutherford fostered and influenced an amazingly gifted group of young men, including Niels Bohr.

"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 fifteen-inch shell at a piece of tissue paper and it came back and hit you." [see Figure 8.8 above]
Ernest Rutherford

the same size as an atom. The small, relatively light electrons were sprinkled throughout this region, somewhat like seeds in a watermelon or raisins in a plum pudding (Fig. 8.6). Since the negatively charged electrons repel one another, they tend to spread apart. The attraction of the positive charge, however, kept them from separating completely.

Thomson concluded that the electrons in this model should arrange themselves in ring-like layers. He also thought that atoms were held together in molecules and crystals by electrical forces between the unbalanced charges that are created when electrons are lost by one atom (which is then electrically positive) and are acquired by another atom (which is then electrically negative). In addition, Thomson tried, only partially successfully, to estimate how the electrons might vibrate so he could predict the frequencies of the emitted electromagnetic radiation.

Rutherford's experiment. An investigation conducted in Ernest Rutherford's laboratory in 1908 yielded a surprising result that could not be interpreted with the Thomson model for the atom. An extremely thin gold foil was bombarded with alpha rays, which are the disintegration products of the radioactive element radium (Fig. 8.7). From their interaction with magnets and electric charges, it was known that alpha rays are positively charged and are about 8000 times as massive as electrons. Most of the alpha rays that impinged on the foil went through without a measurable deflection, but a very small number were deflected by more than 90°, that is, back towards the source. Apparently, the deflected alpha rays hit something that was present here and there in the gold foil

In the Thomson model, the electrons were present "here and there." Could they have deflected the alpha rays? The very low mass of the electrons compared to that of the alpha rays meant that *none* of the alphas should be deflected by more than 90°, or backwards. *All* deflections from electrons should be small and none large (Fig. 8.8 and Figure 8.9a). But Rutherford's many experiments clearly showed that this was not true; therefore, he had to devise his own model for the way the electrons and the positive charge were arranged in the atom.

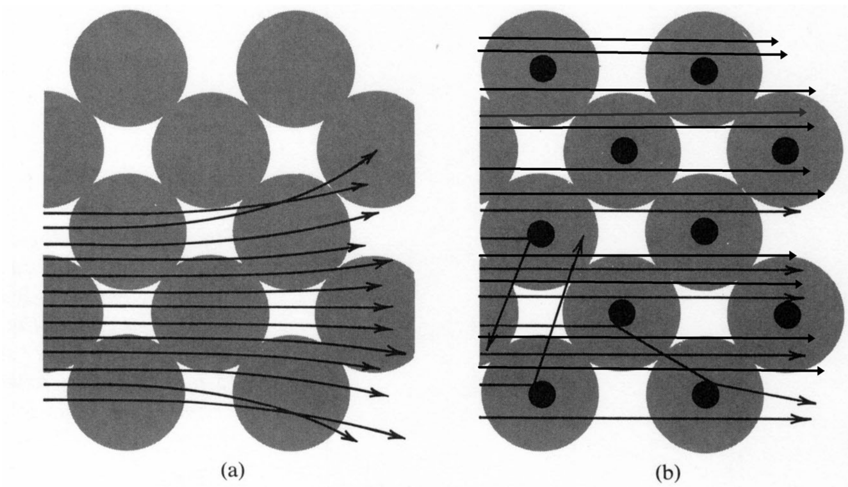


Figure 8.9 Deflection of alpha rays by a gold foil. (a) As predicted by Thomson's "plum pudding model," most alphas would be deflected by small angles, but none would be deflected by more than 90°. (b) As predicted by Rutherford's nuclear model, most alphas would be undeflected, but a few would actually strike the nucleus and thus be deflected by substantial angles. In addition, among alphas actually striking the nucleus, there would be a very few that would have "head-on" collisions. These alphas represent the very small, but not zero, number deflected by more than 90°

Rutherford's nuclear model. If the electrons weren't heavy enough to bounce the alpha rays back, what else in the atom could do it? Rutherford therefore thought further about the positive charge in the atom, which Thomson's model represented as being spread out like a cloud, but which included almost all the mass of the atom. Perhaps it was not spread out; perhaps it was actually concentrated with all its mass in a very small region of space. If alpha rays strike such a small, massive object head on, they would recoil in the direction from which they came or be deflected by large angles, as shown in Figure 8.9b. The alpha rays are like random "bullets" striking a target; only the ones that happen to hit the "bull's-eye" (the positive charge) would be deflected by large angles. Most of the alpha rays would simply continue in a straight line. The percentage of deflected alphas should be directly related to the size of the bull's-eye.

On the basis of this relationship Rutherford estimated the size of "bull's-eye," that is, the region occupied by the positive charge within the atom. Rutherford found that this region was much, much smaller than the size of a gold atom. Rutherford called this small, massive, positive charge the *nucleus* of the atom.

The atomic model invented by Rutherford, called the *nuclear atom,* has stood the test of time and, with many refinements, is still accepted today. However, when Rutherford proposed the nuclear model, there were many serious difficulties and various "mysteries" associated with this new model. We will now explain these problems and how they were resolved, which laid the groundwork for quantum theory.

Difficulties of the nuclear model. One of the mysteries was to explain how all the positive electric charge could remain clustered in a small region of space in spite of the enormous mutual repulsion of all the parts of the charge for one another. To solve this problem, it is necessary to make a model for the atomic nucleus itself, showing the parts it is composed of and how they interact (Section 8.6).

The problem of atomic stability. A second mystery was why the negative electrons, which were strongly attracted by the positive nucleus, remained outside the nucleus and were not pulled into it. The first attempt to resolve this mystery was not successful. Rutherford suggested that the atom is like a miniature solar system, with the light, negative electrons orbiting around the heavy, positive nucleus: the planetary model for the atom. Even though such an atom would not collapse immediately, the relative motion of negative and positive charges was known to lead to the emission of electromagnetic waves, as in Hertz's spark gap experiments (Section 7.3). These waves should be observable as light emitted by atoms and would gradually rob the atoms of energy until the electrons eventually collapsed into the nucleus (Fig. 8.10). Since atoms are not observed to emit light until they collapse (in fact, they are stable and do not collapse), the planetary version of the nuclear atom did not seem satisfactory.

The problem of line spectra. A third mystery was the line spectrum of light emitted by hot gases (Section 7.2). Hot gases emit light of selected frequencies, and the atomic models should give an explanation of the spectrum. The planetary nuclear atom led to the emission of light, as we have just explained, but the emission never stopped, and the frequency varied as the electron spiraled with an increasing frequency on its way toward the nucleus.

In spite of its difficulties, the nuclear atom also enjoyed successes. By studying the deflected alpha rays, Rutherford was able to determine the relative amounts of positive electric charge on various nuclei. He found that the positive charges on nuclei of different elements, such as carbon, aluminum, copper, and gold, varied in the same ratio as the atomic number of the element (Section 4.5). Since the negatively charged electrons in an atom must neutralize the positive charge on the nucleus, the number of electrons is just equal to the atomic number and varied from one element to the next in the periodic table. This startling result explained how the atoms of various elements differed from one another; furthermore, it explained the significance of the chemists' atomic number in terms of the structure of atoms.

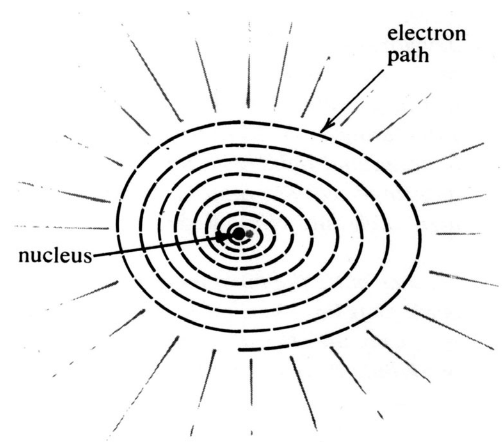


Figure 8.10 In the planetary model, the electron emits light and spirals into the nucleus as it loses more and more energy.

8.3 Bohr's model for the atom

Bohr's quantum rules:
1. Only certain selected ones of all conceivable orbits of the planetary motion of the electrons around the nucleus are permissible.
2. When an atom emits (or absorbs) light, it makes a quantum jump from one to another of the selected orbits. The energy transferred, called a quantum of energy, determines the frequency of the light, according to Eq. 8.1.

Equation 8.1

energy transferred = ΔE
frequency of light = f
Planck's constant = h

$$\Delta E = hf$$

Planck's constant h has a numerical value in the micro domain,

$$h = 6.6 \times 10^{-34}\ \text{kg-m}^2/\text{sec}$$

It therefore does not influence macro-domain phenomena directly.

" ... an electron of great velocity in passing through an atom and colliding with the [bound] electrons will lose energy in distinct finite quanta ... very different from what we might expect if the result of the collisions was governed by the usual mechanical laws."

Niels Bohr
Philosophical Magazine, 1913

Bohr's theory. Niels Bohr, a young Danish theoretical physicist temporarily working at Rutherford's laboratory, initiated a completely new approach to atomic models and achieved a decisive breakthrough in 1912. He accepted the findings of Thomson and Rutherford, but he did not submit their models merely to the laws of motion and electromagnetism as formulated by Newton and Maxwell, respectively. In addition, he was stimulated by the inconsistencies of the existing models to introduce new principles called *quantum rules* (stated to the left), and he found that these solved all the difficulties associated with the electron motion and light emission in the nuclear atom. Equation 8.1 is one of the quantum rules, called *Bohr's frequency condition.* The quantum rules and the procedures for using them make up Bohr's theory. When Bohr applied the quantum rules to the nuclear model for the hydrogen atom, which is the simplest of all atoms in that it has only one electron, he obtained excellent quantitative agreement with observations of the hydrogen spectrum. This explanation of the frequencies of the light emitted and absorbed by the hydrogen atom was a startling success of Bohr's theory.

A little later we will explain the details of the quantum rules and how to apply them. Now we will describe how they eliminate the critical difficulties of the nuclear model.

States of an atom. In any model for atoms, the state of an atom is described in terms of the arrangement and motion of its parts. Before Bohr, the radius of an electron's orbit around the nucleus was believed to become smaller and smaller as the electron lost energy by radiation and spiraled into the nucleus. The energy of the atom in any state was calculated from the speeds of the electron (kinetic energy) and the electric interaction among the electrons and the nucleus (electric field energy). With Bohr's rules, certain electron orbits and therefore certain states of atoms were selected, and all others were eliminated. The energy of each allowed state, however, was calculated from the kinetic and electric field energies as before. Each allowed state of an atom is called an *energy level* (Fig. 8.11).

Two important consequences follow from the existence of energy levels. First, there is an atomic state with the lowest permitted energy, called the *ground state.* An atom in the ground state cannot emit light, because emission of light would reduce its energy below the lowest permitted value. The ground state is, therefore, safe against collapse of the electrons into the nucleus. Second, when the atom is not in the ground state, it can emit light as the electron "jumps" from an orbit of higher energy to another of lower energy. The energy transferred to the light is a specific amount equal to the energy difference between the two states of the atom (Fig. 8.12). The frequency of the light can be calculated from the energy transfer according to the formula in Bohr's second rule and his frequency condition (Eq. 8.1). Hence the emission spectrum will consist only of the spectral lines arising from all the allowed orbital "jumps" of the electrons; the spectrum will not be a complete rainbow including all frequencies.

Niels Bohr (1885-1962) had just received his doctorate in Denmark when he went to England in 1911. He joined Rutherford's research group in Manchester, the "home" of the nuclear model of the atom. Bohr was intimately familiar with the latest physics theory (mostly developed in continental Europe, especially Germany, England's archrival) which revealed the stubborn contradictions and difficulties inherent in the nuclear model, and he had a remarkable capacity to connect with others as well as boundless enthusiasm for discussing the details of apparently conflicting ideas. The 26 year-old Danish theorist responded to the stimulation of Rutherford's lab with an extraordinary burst of creativity: audaciously inventing his own new "quantum rules" and using them, together with the latest theory, to resolve the well-known problems delaying acceptance of the nuclear atom. For this work, published in 1913 and 1915 (while Europe was entering World War I), Bohr was awarded the Nobel Prize in 1922.

Bohr's scientific genius was equaled by his courage, compassion, and wisdom. During World War II he helped rescue thousands of Jewish refugees by smuggling them out of Nazi-occupied Denmark to the safety of Sweden. Bohr's last years as Director of the Copenhagen Institute for Theoretical Physics were distinguished by unceasing efforts to secure international peace in an age threatened by nuclear holocaust.

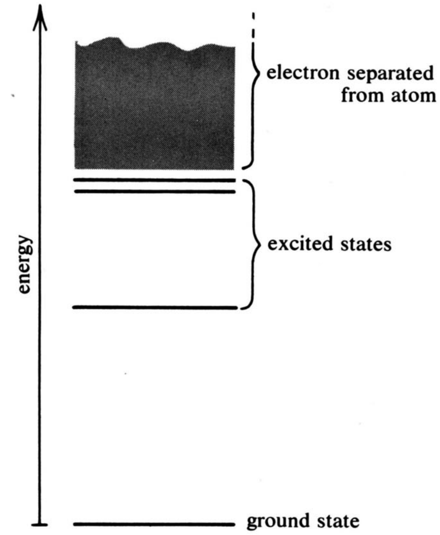


Figure 8.11 Diagram to represent the allowed states of an atom. Each line represents one state allowed by Bohr's quantum rules. The distance between lines represents the energy difference between the states. The diagram is called an energy level diagram.

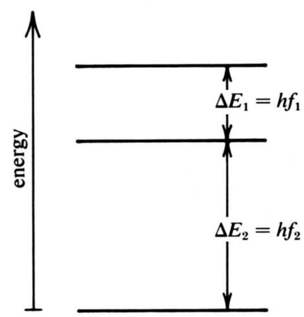


Figure 8.12 The frequencies of emitted spectral lines are found from the energy differences on an energy level diagram.

Revolutionary aspect of Bohr's rules. Bohr's theory provided a breakthrough because it added to the accepted laws and principles of physics. It did not merely combine the accepted laws to obtain a new conclusion. This is characteristic of a scientific revolution.

Bohr's rules do seem arbitrary and ad hoc, and they appear to conflict with other, well-established laws of physics. The question naturally arises why should Bohr's rules hold? In 1912 many physicists asked this question. However, a "why" question of this type is more a philosophical than a scientific question. The quantum rules are an integral part of Bohr's model: the rules must be accepted if the model is successful and rejected, or modified, if it fails. If a second successful model is available, we could choose the one that better satisfies our curiosity or preferences. But, in 1912, calculations based on Bohr's theory, and no other, produced sensible results that agreed precisely with the observed frequencies of light emitted and absorbed by the hydrogen atom. With only one model, we do not have a choice. Nevertheless, many of Bohr's contemporaries questioned the new theory, and several years passed before the quantum rules were accepted for their value.

Quantum ideas before Bohr. The idea of the quantum (and the associated idea that micro-domain systems had discrete, not continuous,

Equation 8.2

energy of oscillator = E
frequency of oscillator = f
Planck's constant = h

$E = hf$

"While this constant was absolutely indispensible to the attainment of a correct expression ... it obstinately withstood all attempts at fitting it, in any suitable form, into the frame of a classical theory."

Max Planck
Nobel Prize address, 1920

Equation 8.3

energy of 1 quantum = E

$E = hf$

Equation 8.4

electron momentum = $\mathcal{M}$
orbit circumference = $2\pi r$
quantum number = n

$2\pi r \mathcal{M} = nh$

$(n = 1, 2, 3, 4 \ldots)$

Equation 8.4 restricts the radius and momentum of the electron to only certain allowed values. The radius r and the momentum $\mathcal{M}$ must have values that fit into one of the following:

$2\pi r\mathcal{M} = h$, or
$2\pi r\mathcal{M} = 2h$, or
$2\pi r\mathcal{M} = 3h$, and so on….

energy levels, as shown in Figure 8.12) did not originate with Bohr. Max Planck (1858-1947) had introduced it into the theory of light emission from glowing solids (Section 7.2). Planck used the micro-domain model that represented a solid as a system of electrically charged oscillators (Section 7-3). To explain the data, Planck assumed that only selected states of the oscillators could radiate their energy and that the energies of these states were directly proportional to the oscillator frequency (Eq. 8.2). The formula relating the energy to the frequency depended on a new number, which Planck called the *quantum of action* (symbol *h*). This number (now called Planck's constant) and formula were similar to the ones introduced later by Bohr.

A few years after Planck's work, Einstein applied the quantum concept to the photoelectric effect (Fig. 8.13). When light is absorbed by certain metals, electrons may be ejected from the metal surface. The frequency of the light and the kinetic energy of the electrons were measured and compared. Below a certain frequency, no electrons were emitted. Above that frequency, the kinetic energy increased in a regular fashion when the light frequency was increased. Einstein found he could explain the experimental measurements in terms of energy transfer from the light to each ejected electron with the following assumption: radiant energy can be transferred to one electron only in certain definite amounts, called a quantum. The energy of one quantum of light is directly proportional to its frequency, with Planck's constant again making its appearance (Eq. 8.3).

Quantum number—Applications of Bohr's quantum rules. Bohr used Planck's constant and the electron momentum for the application of his first quantum rule to specific cases. You will recall that the momentum of a particle is its inertial mass multiplied by its speed (Eq. 3.1). The allowed circular orbits (quantum states) of Bohr's first rule are selected with the requirement that the electron momentum multiplied by the circumference of the orbit is equal to a whole number times Planck's constant (Fig. 8.14 and Eq. 8.4). The whole number n that appears in Eq. 8.4 is called a quantum number. This is known as quantization in which some quantity (in this case, $2\pi r\mathcal{M}$) can have only certain restricted values that are multiples of a basic quantity (in this case h).

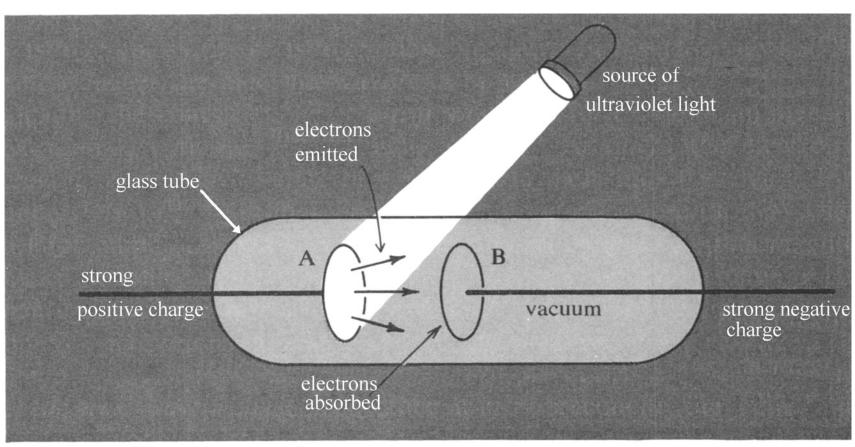


Figure 8.13 (to the right) The photoelectric effect. When ultraviolet light illuminates copper plate A, electrons are emitted by the copper plate, leaving it with a positive electric charge. The electrons are absorbed by copper plate B, giving it a negative electric charge.

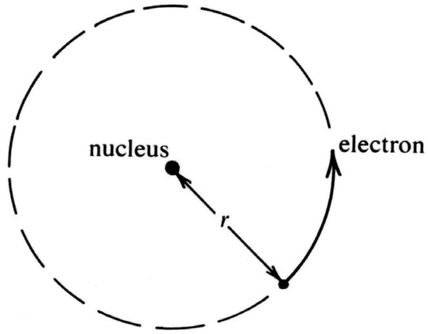


Figure 8.14 Electron in circular orbit around the nucleus. The radius of the orbit is r, the circumference is 2πr.

In Bohr's second rule, the energy transferred when an electron jumps from one orbit to another is directly proportional to the light frequency (in (Eq. 8.1) in the same relation with Planck's constant as found by Planck and Einstein (Eqs. 8.2 and 8.3). The quantity of energy transferred can be expressed by means of the quantum numbers of the two orbits. As we stated above, this model predicts quite accurately the observed spectrum of the hydrogen atom. This outstanding success of Bohr's theory made it very attractive to many physicists in spite of its revolutionary nature, and intensive research to test its implications was undertaken.

Limitations of Bohr's theory. With the passage of time and accumulation of new, more precise data, Bohr's theory had to be modified in minor ways. However, more serious inadequacies were revealed in its applications to complex atoms containing many electrons. Bohr's theory led to the conclusion that all the electrons in the atom's ground state would occupy the same orbit close to the nucleus. Both the spectra and the sizes of atoms, however, indicated that this did not happen, but that electrons arranged themselves in different orbits, some close to the nucleus, some farther away. Furthermore, the repulsive interaction of the electrons with one another could not be included satisfactorily in the theory. The difficulty of treating systems containing many electrons made it impossible to apply the model to the formation of molecules out of atoms, because all molecules do include many electrons.

In addition, the model suffered from the philosophical weakness that it combined the old laws of Newton and Maxwell with the new quantum rules. This rather capricious combination was unattractive for some physicists, who could not see clearly just when the old and when the new was to be employed. Nevertheless, Bohr's theory was the first step, the breakthrough, of a scientific revolution in the theories and models for micro-domain physical systems. This revolution, which reached its climax in the nineteen twenties, was throughout inspired and guided by Bohr.

8.4 The wave mechanical atom

In Bohr's mysterious quantization, the electron momentum multiplied by the orbit's circumference is equal to a whole number times Planck's constant (Eq. 8.4). Another way to state this condition is that the length of the path of the electron in its orbit (2πr) is a whole number times Planck's constant divided by the electron momentum (Eq. 8.5)

Equation 8.5

$$2 \pi r = n (h / \mathcal{M})$$

$(n = 1, 2, 3, ...)$

Electron waves. Now, this relation reminded the French physicist Louis de Broglie of the condition for the wavelength of waves permitted in a tuned system (Eq. 8.6; see Eq. 6.4). Using Bohr's model for the atom, de Broglie interpreted the circumference of the orbit as the length of the tuned system. By comparing Eqs. 8.5 and 8.6 and identifying the circumference, 2πr, with L, we can see the close similarity between $h/\mathcal{M}$ and the wavelength, λ. De Broglie therefore introduced the idea that the motion of electrons is governed by a wave whose wavelength is inversely proportional to the electron momentum (Eq. 8.7a) and whose wave number (k = 1/λ) is therefore directly proportional to the momentum (Eq. 8.7b). An atom is thus viewed as a tuned system that contains standing waves of one or more electrons. The electrons are refracted around the positively charged nucleus by its electrical field, which affects electron waves in the same way that a change of index of refraction affects light waves (Fig. 8.5).

Equation 8.6

length of tuned system = L
wavelength = λ
number of waves = n

$$L = n\lambda$$

Equation 8.7

wave number = k

$$\lambda = h / \mathcal{M} \quad \text{(a)}$$

$$k = \mathcal{M} / h \quad \text{(b)}$$

Free electrons. If the motion of electrons is really governed by waves while they orbit the nucleus, the same should be true for electrons in a cathode-ray beam. Such electrons should therefore be diffracted by a suitable "grating" whose slits are separated by a distance comparable to the wavelength as given in Eq. 6.13. A nickel crystal provides such a grating. And, as we pointed out in Section 8.1, Davisson and Germer accurately verified this prediction of de Broglie's model.

Wave mechanics. The branch of physics in which the motion of electrons is represented as wave propagation is called wave mechanics. Electron waves are diffracted by obstacles. They are refracted when passing through the electric fields of the nucleus and of other electrons. The computational method for finding electron standing wave patterns was developed by the Austrian physicist Erwin Schrödinger and has been successfully applied to simple atoms and molecules. In principle, wave mechanical calculations make it possible to compute the physical and chemical properties of all substances. In practice, the calculations could be done for only the simplest atoms. At that time, the calculations for atoms with more than 2 or 3 electrons were much too complicated; now high-speed computers have successfully calculated the properties

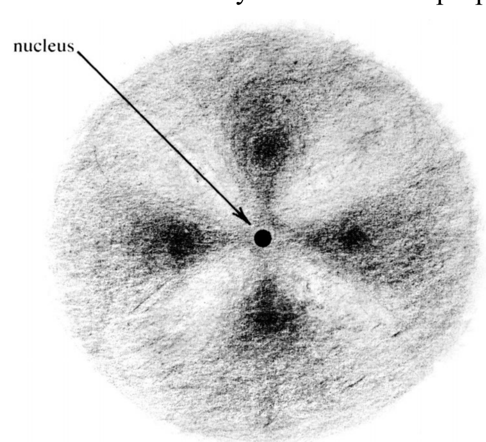


Figure 8.15 Diagram showing electron waves refracted around the nucleus. Four wave maxima and four minima are shown. Compare with Fig. 8.14, which represents the orbit of electron particles.

of many atoms and molecules; thus confirming the validity of the wave mechanical model to extremely high accuracy.

Electron waves in atoms and molecules. According to the wave mechanical model, therefore, an atom consists of a small nucleus surrounded by a standing wave of one or more electrons. The atom is very much larger than the nucleus, with most of its space occupied by the electron waves. The size of the atom is approximately equal to one or a few wavelengths of the electron waves forming a standing wave pattern. The atomic collapse predicted in the particle model for the electron does not occur because the particle description is not applicable. The separation of adjacent atoms in a molecule or crystal is maintained by the mutual repulsion of the negatively charged electron waves surrounding the nuclei of adjacent atoms. The formation of molecules is explained by the partial merging of certain electron waves, called valence electrons, for which the attraction by the adjacent nucleus is stronger than the repulsion by the adjacent electron wave.

Interaction of atoms with light. The wave mechanical model retains the concept of quantum jumps in which the electron gains or loses energy when the atom absorbs or emits light. In a quantum jump, the electron standing wave patterns are transformed from those of the initial state to those of the final state. The frequency of the light, still determined by Bohr's frequency condition (Eq. 8.1), is directly proportional to the energy gain or loss. Schrödinger's mathematical formulation, however, makes it possible not only to calculate the frequency but also to relate the intensity of the emitted light to the standing wave patterns in the atom's initial and final states.

Louis Victor de Broglie (1892-1987) is unique among great physicists because he became interested in physics at an age when most physicists have completed their major work, and even this late start was interrupted by service in the army during World War I. After the war, in his brother's scientific laboratory, the idea that electrons may be considered to have wave properties occurred to him. This theory, delivered in his doctoral thesis, Recherches sur la Theorie des Quanta, in 1924, won de Broglie the Nobel Prize in 1929.

Erwin Schrödinger (1887-1960) was born and educated in Vienna. Violently anti-Nazi, Schrödinger barely escaped a concentration camp by fleeing Austria in 1940. Most of his last years were spent as director of the Institute for Advanced Studies in Dublin, Ireland, where his interest was attracted to questions related to the nature of life. In 1924 and 1925 Schrödinger developed wave mechanics using de Broglie's idea that stationary states in atoms correspond to standing matter waves. For this distinguished work, Schrödinger shared the Nobel Prize with P. A. M. Dirac in 1933.

8.5 Wave particle duality and the uncertainty principle

You may wonder how it is possible to reconcile the particle and wave models for the electron. Cathode rays were originally invented as the intermediaries in an interaction-at-a-distance between the cathode and the glass of the discharge tube. When it was found that they possessed electric charge and inertial mass, a particle model became accepted; the particles were called electrons. Later, the diffraction experiments gave evidence of the electron's wave nature.

The answer to the question "Where is the electric charge of the electron located?" illustrates the apparent contradictions between the two models. In the particle model, the charge is located at the position of the particle. In the wave model, the charge is spread throughout the region occupied by the wave; the electron is found more often (and the charge is more concentrated) where the wave has a large amplitude and the electron is found less often (and the charge is less concentrated) where the wave has a small amplitude. It should be possible, you may say, to find the correct model by means of an experiment. Unfortunately, as we will explain below, the experiments that can be carried out do not help to distinguish between the two models. As a matter of fact, there was a famous years-long debate between Einstein and Bohr, in

The letters below illustrate the long-running debate between Bohr and Einstein about the probability interpretation of quantum theory. In 1949 (when quantum theory, including the probability interpretation, had been well accepted by essentially the entire physics community) Bohr sent a letter to Einstein wishing him a happy 70th birthday.

Einstein wrote back, thanking Bohr for his good wishes and added, "This is one of the occasions, which is not dependent on the disquieting question if God throws the dice or if we should hold on to the available physical description of realities."

Bohr then wrote back, ". . .To continue in the same jocular tone, I have no choice but to say on this painful issue that it really doesn't matter in my opinion whether or not we should hold on to the physical description of accessible realities or further pursue the path you have shown and recognize the logical assumptions for the description of realities. In my impertinent way I would suggest that nobody—not even God himself—would know what a phrase like playing dice would mean in that context."

"Like following life through creatures you dissect, You lose it in the moment you detect."
Alexander Pope
Essay on Man, 1732

which Einstein tried to find weaknesses in the quantum theory, but in which (in the opinion of most physicists today) Bohr actually found the weaknesses of Einstein's objections.

The effect of measurement on the state of a system. Consider the following simple example. A blind person tries to find the location of a glass marble on a table. He feels with his hands, measuring from the edges of the table, until his fingers touch the marble. Then he knows where he found it, but he will also have bumped it slightly and nudged it to a new location.

In any real experiment, some "nudging" of the system under study always takes place. What we tend to do, however, is to idealize by imagining a very, very gentle "nudge"—just enough, as it were, to tell the blind person that he has found the marble, but not enough to displace it. You can readily see that this ideal does not really exist, that finding the marble will always influence the marble, and observing a system will always influence or disturb the system in some way.

Even though the scientist has much more delicate instruments than the blind man's hands, the quantization of energy limits the "gentleness" with which he can operate. When he detects the position of the electron, there has to be some interaction and energy transfer between the electron and the measuring instrument. The smallest amount of energy that can be transferred is 1 quantum. Since the electron's inertial mass is very small, even the transfer of a very small amount of energy greatly affects its state.

Probability interpretation. The structure of the quantum theory, in other words, makes it impossible to carry out operations that identify an electron definitely as a wave or a particle. When you treat the electron as a particle and direct it through a small opening, the resulting diffraction pattern reminds you of its wave nature. When you treat the electron as a wave and try to separate a portion of the wave pattern from the remainder, you find that you get either all of it or none of it, as you would for a particle. The question "Is the electron a wave or a particle?" therefore does not have an operational meaning within the theory.

Nevertheless, there is a relationship between the two views. The amplitude of the electron wave can be related to the probability for locating the charge of the electron. Suppose, for example, an investigator makes many measurements of the position of the electron's charge in the ground state of an atom. He then finds the charge frequently (that is, with high probability) in regions where the wave has a large amplitude (Fig. 8.15). He finds the charge rarely (that is, with low probability) in regions where the wave has a small amplitude. He has to make many observations, however, and each time he finds all or none of the charge of the electron at the position he is observing.

Electron wave packets. Another way to reconcile the wave and particle models is to represent the electron as a wave packet. A wave

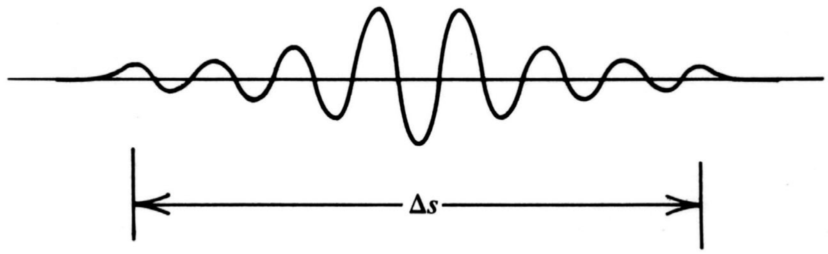


Figure 8 16 An electron wave packet of size As.

Equation 8.8

size of the wave packet
$= \Delta s$
uncertainty in wave number
$= \Delta k$

$$\Delta s \, \Delta k = 1$$

Equation 8.9

uncertainty in wave number
$= \Delta k$
uncertainty in momentum
$= \Delta \mathcal{M}$

$$\Delta k = \Delta \mathcal{M} / h$$

Equation 8.10

$$\Delta s \, \Delta \mathcal{M} = h$$

You must remember that a particle is an idealized object used in the construction of models. A particle occupies a point in space and has inertial mass, speed, and momentum. A wave packet is an alternative, somewhat more complex, idealized object which we use instead of a particle in constructing models in the micro domain.

packet differs from a single wave train in that it does not extend throughout space. It differs from a particle in that it is not localized at one point. Wave packets are described by an uncertainty principle (Eq. 8.8), which was explained in Section 6.2. This principle applies to electron waves also.

We will now interpret the uncertainty principle for electron wave packets. The size Δs of the wave packet (Fig. 8.16) is interpreted as the uncertainty in the position of the electron. The uncertainty Δk in the wave number is interpreted as an uncertainty in the momentum of the electron (Eq. 8.9, derived from de Broglie's relation Eq. 8.7b). You can therefore infer an uncertainty relation between the position and the momentum of the electron (Eq. 8.10). The smaller the uncertainty in the position, the larger the uncertainty in momentum. And vice versa. It is not possible to find the precise position and momentum of an electron simultaneously. This property of an electron means that it cannot be described by a particle model.

Matter waves. Experiments with beams of alpha rays, atomic nuclei, and even entire atoms and molecules have shown that these, too, are diffracted; in other words, their motion is in accord with the theory of wave propagation. You may now ask, "Where does this end? Does all matter exhibit wave properties?" The answer is, in principle, yes. A wave packet is associated with every material system. In practice, however, the wavelengths of systems in the macro domain are so small that diffraction effects are undetectable for them. The uncertainty principle, for example, reveals that a speck of dust with a mass of 1 milligram can be treated as a "particle" with negligible uncertainty of position and momentum, even though an electron cannot be treated as such (Example 8.1). The reason is the micro-domain numerical value of Planck's constant.

EXAMPLE 8·1 Applications of the uncertainty principle using Eq. 8·10, $\Delta s \approx h/\Delta\mathcal{M}$.

(a) Electron:
mass $M_I = 10^{-30}$ kg
speed $v = 10^{7}$ m/sec

momentum $\mathcal{M} = M_I \, v = 10^{-30}\text{ kg} \times 10^{7}\text{ m/sec} = 10^{-23}\text{ kg-m/sec}$

ten percent uncertainty in momentum:

$$\Delta\mathscr{M} = 10^{-24} \text{ kg-m/sec}$$

uncertainty in position:

$$\Delta s \approx \frac{h}{\Delta\mathscr{M}} = \frac{6.6 \times 10^{-34} \text{ kg-m}^2\text{/sec}}{10^{-24} \text{ kg-m/sec}} = 6.6 \times 10^{-10} \text{ m}$$

(approximately five times the diameter of a hydrogen atom)

(b) Dust particle:
mass $M_I = 10^{-6}$ kg
speed $v = 1$ m/sec

$$\text{momentum } \mathscr{M} = M_I v = 10^{-6} \text{ kg-m/sec}$$

one-tenth percent uncertainty in momentum:

$$\Delta\mathscr{M} = 10^{-9} \text{ kg-m/sec}$$

uncertainty in position:

$$\Delta s = \frac{h}{\Delta\mathscr{M}} = \frac{6.6 \times 10^{-34} \text{ kg-m}^2\text{/sec}}{10^{-9} \text{ kg-m/sec}} = 6.6 \times 10^{-25} \text{ m}$$

(a completely negligible uncertainty)

8.6 The atomic nucleus

In the preceding section, the atomic nucleus was described as a massive, positively charged particle whose electric field refracted the electrons into a standing wave around the nucleus. As a matter of fact, scientists have formulated models for the nucleus itself as a complex system of interacting parts, a system capable of acting as energy source or energy receiver by changing its state.

Radioactivity. Already before Rutherford's invention of the nuclear atom, mysterious rays (including Rutherford's alpha rays) had been observed by Henri Becquerel to affect photographic plates near certain materials, especially compounds containing uranium, and had been given the name radioactivity. These rays were soon identified by their interaction with a magnetic field as having three components (Fig. 8.17): massive, positively charged alpha rays (Rutherford's tool for discovering the nucleus); light, negatively charged beta rays (later identified as electrons); and electrically uncharged gamma rays, which acted similarly to X rays (later identified as electromagnetic radiation). Scientists studied the rays and the sources intensively, finding that certain elements must be the source of the rays and that, amazingly enough, these elements seemed to be "decaying" into new elements (the objective of the alchemists' centuries-old, but unsuccessful, quest). Initially, because the rate of all known chemical reactions depended on temperature, it also seemed obvious that the rate at which the rays were emitted should depend on temperature, or on the specific other elements with which the radioactive elements were combined. However, many studies established conclusively the opposite: the rate of radioactivity depended only on the amount of the particular radioactive elements present, and neither on the temperature nor anything else.

Radioactive elements each have a characteristic half-life (the time it takes for half of a sample of the element to decay into another element, leaving one half of the original element). The half-life of radium is about 1600 years, and uranium's is almost five billion years! On the other hand, the half-life of other radioactive elements is a tiny fraction of a second.

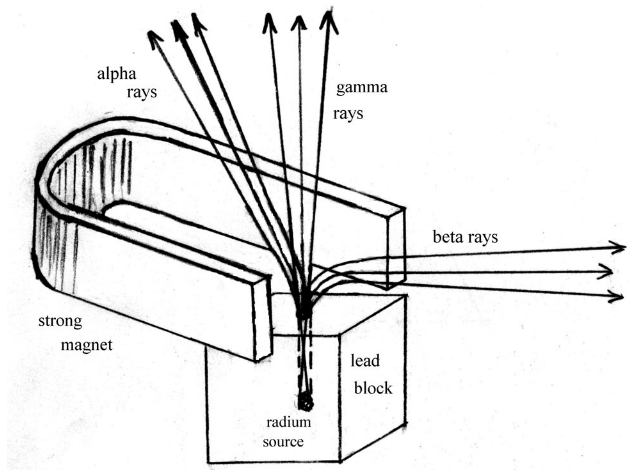


Figure 8.17 Becquerel's separation of alpha, beta, and gamma rays by using a magnetic field. Compare the deflections with those of cathode rays in Figure 8.4.

Henri Becquerel (1852-1908) studied at the Ecole Polytechnique, where he was appointed a demonstrator in 1875 and a professor in 1895. His accidental discovery of radioactivity in 1896 triggered an avalanche of research into radioactivity and its source in the atomic nucleus.

" I particularly insist on the following fact, which appears to me exceedingly important and not in accord with the phenomena which one might expect to observe: the same encrusted crystals placed with respect to the photographic plates in the same conditions and acting through the same screens, but protected from the excitation of incident rays and kept in the dark, still produce the same photographic effects."
Henri Becquerel
Comptes Rendus, 1896

The discovery of helium in radioactive ores and the experimental similarity of alpha rays to electrically charged helium atoms led Rutherford in 1903 to the conclusion that radioactivity was a breakdown of an atom of one element into an atom of another element. Uranium was gradually converted to radium and finally to lead, with alpha rays being emitted. The alpha rays then form helium atoms, which accumulate in the material.

This was the first serious attack on the concept that atoms were indivisible and immutable. (The next major step was Rutherford's invention of the nuclear atom.) The electric charges, masses, and energies observed in radioactivity quickly led Rutherford to the further conviction that radioactivity was a breakdown of the nucleus itself and did not involve the atomic electrons in a significant way. From the very beginning of the nuclear model it was clear that the nucleus was divisible and changeable. It was natural, therefore, for Rutherford to be puzzled that the positive charges in the nucleus held together as well as they did. Because of the mutual electrical repulsion of the positive charges, he wondered why every atomic nucleus did not disintegrate, as radioactive nuclei in fact do when they emit alpha rays.

Modern nuclear models. According to the currently accepted model, the atomic nucleus consists of positively charged constituents called *protons* and electrically neutral constituents called *neutrons,* all refracted into a standing wave pattern by their interaction. Neutrons and protons have approximately the same mass, but each is about 2000 times as massive as an electron. The nucleus of a hydrogen atom is one proton, an alpha ray consists of two protons and two neutrons, an oxygen nucleus includes eight protons and eight neutrons, and a uranium nucleus contains 92 protons and 143 to 146 neutrons.

Atomic number. The positive electric charge of the proton is

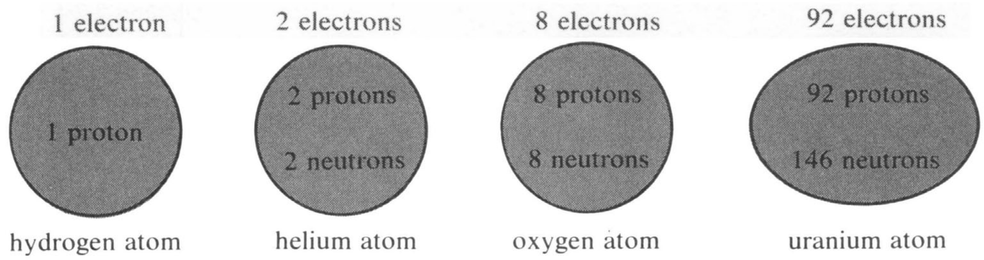


Figure 8.18 (above) The composition of four electrically neutral atoms. The numbers of electrons and protons are always equal.

equal in magnitude to the negative electric charge of the electron. Therefore, the hydrogen atom composed of one proton and one electron is electrically neutral. The number of protons determines the total positive electric charge of the nucleus, and this, in turn, determines the number of electrons that are bound in the neutral atom (Fig. 8.18). The number of electrons, finally, determines the chemical properties of an element (Section 8.2). The chemical properties, in this model, are therefore traceable to the number of protons in an atomic nucleus of the element. The number of neutrons in the nucleus does not play a role in an element's chemical properties but it does contribute to the mass of the nucleus and to its stability.

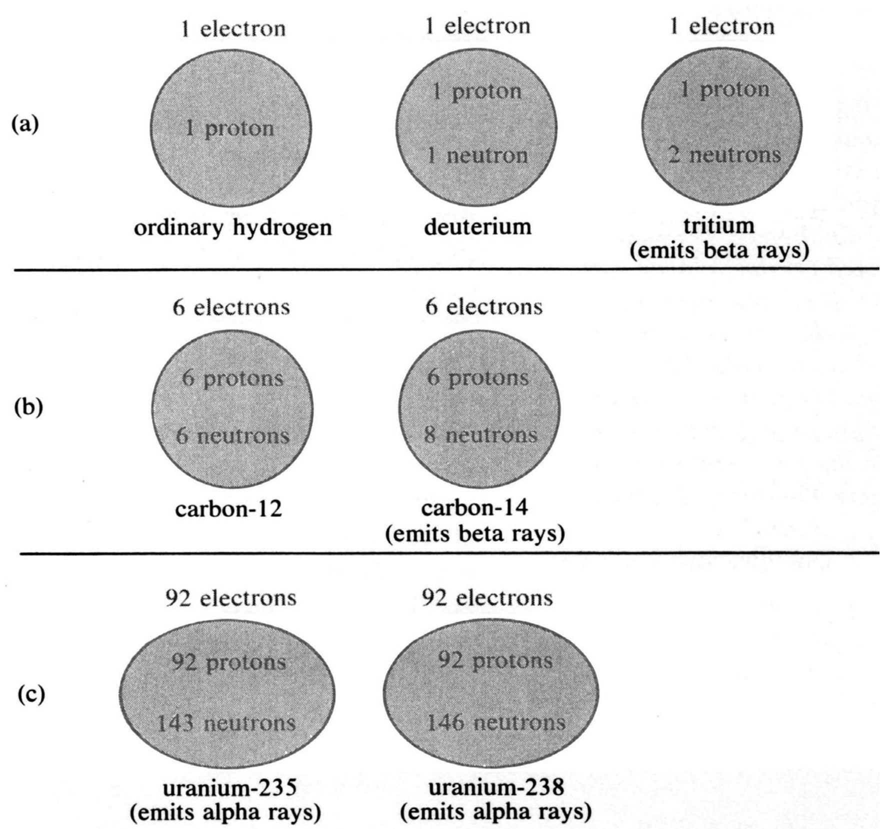
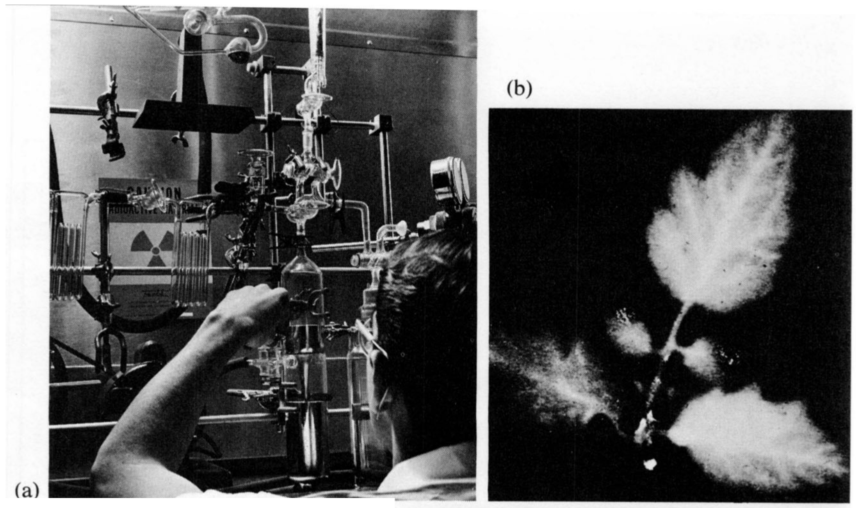


Figure 8.19 Diagrammatic representation of isotopes.
(a) Isotopes of hydrogen.
(b) Isotopes of carbon.
(c) Isotopes of uranium.



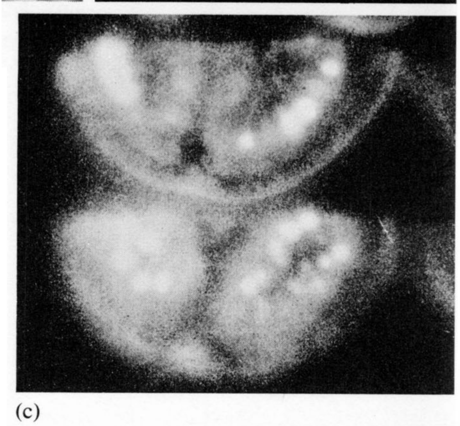


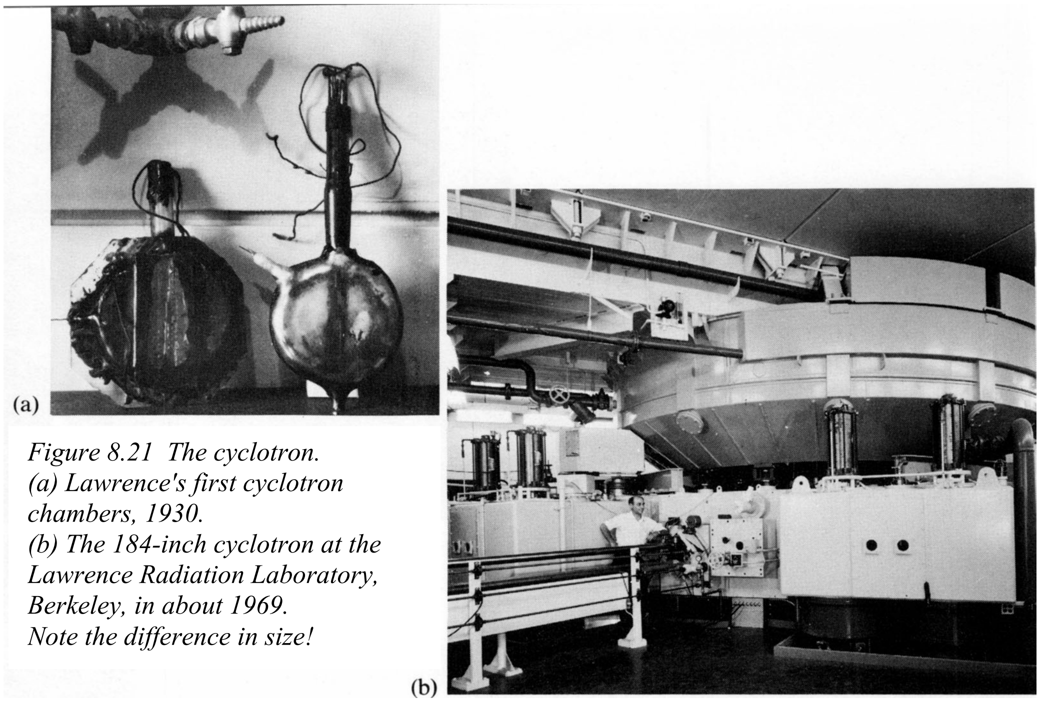
Figure 8.20 Application of radioactive isotopes.
(a) Radioactive carbon is being used in this drug metabolism study. Drugs labeled with radioactive materials yield information about concentration of compounds in tissue and about molecular structures, and provide clues to biosynthesis and interactions of normal body chemicals.
(b) The first "radio autograph" made in the dark with an artificial isotope of phosphorus. The isotope, produced in E. O. Lawrence's cyclotron, shows its presence in the leaf by affecting the photographic plate on which the leaf was placed.
(c) Radio autograph of a sliced tomato fed with a trace of radioactive zinc. The lightest spots show where the radioactivity is strongest. What can you infer about the role of zinc in growing tomatoes?

Isotopes. The name "isotope" is given to atoms of one element whose nuclei differ only in their number of neutrons (Fig. 8.19). Most elements have one or a few known stable isotopes and a few radioactive isotopes, which emit beta rays. The radioactive isotopes are thereby transformed into stable isotopes of another element. Isotopes have found many uses in industry, science, and medicine (Fig. 8.20).

In the high temperatures at the center of stars atomic nuclei become separated from their electrons. Thus, the nuclei are not surrounded by mutually repelling electron waves; therefore, the nuclei can interact with each other when they collide. In fact, such collisions result in the large nuclear energy release that, in turn, maintains the high temperature of the sun and stars.

Nuclear stability. Like electrons, both neutrons and protons exhibit wave-particle duality and obey the uncertainty principle. Most important, protons and neutrons participate in a non-electrical attractive interaction-at-a-distance, which binds them to form stable nuclei. The interaction, called the *nuclear interaction,* is much stronger than the electrical repulsion for inter-particle distances less than about 10^{-15} meters, but it becomes exceedingly weak for larger distances. The nuclear interaction acts over such extremely small distances that not even the nuclei of adjacent atoms in solids or liquids are affected. Consequently, there are no macro-domain manifestations of the nuclear interaction and it does not play a role in everyday phenomena.

The role of the neutrons in the nucleus appears to be one of stabilizing the nuclear system. They participate in the strong attractive



(a)

(b)

Figure 8.21 The cyclotron. (a) Lawrence's first cyclotron chambers, 1930. (b) The 184-inch cyclotron at the Lawrence Radiation Laboratory, Berkeley, in about 1969. Note the difference in size!

nuclear interaction and not in the mutual electrical repulsion of the protons. Stable nuclei contain a number of neutrons about equal to or somewhat larger than the number of protons (Figs. 8.19 and 8.20). Nuclei that deviate from the ideal are radioactive and disintegrate by the emission of alpha or beta rays to form more stable nuclei.

Nuclear reactions. *The accelerator* is a research tool that has made possible the systematic study of atomic nuclei and their properties. An accelerator produces a beam of protons, electrons, or alpha rays with very high kinetic energy. The first of these machines was the *cyclotron* invented by E. O. Lawrence in 1931 (Fig. 8.21). Since then, new families of accelerators have been designed and built to study the properties of nuclei during highly energetic collisions.

Ernest O. Lawrence (1902-1958), the father of the modern accelerator, was professor of physics at the University of California at Berkeley and the first director of the university's famed Radiation Laboratory. For his researches in atomic structure, development of the cyclotron, and its use in artificially induced radioactivity, Lawrence won the Nobel Prize in 1939. During World War II, he was one of the chief participants in the race to develop the atomic bomb before the Germans

When the energetic rays produced by an accelerator interact with the nuclei in a target such as thin aluminum foil, they set off a series of changes in the state of the target nucleus. These changes are called *nuclear reactions* (analogous to chemical reactions). The result is the formation of new nuclei with a changed number of neutrons and protons. Frequently these nuclei are radioactive isotopes and are useful in scientific research and in medicine.

Energy transfer. One of the most important practical outcomes of the study of radioactivity and nuclear reactions was the discovery that nuclear transformations involve very a large energy transfer from nuclear field energy to kinetic energy of the reaction products. This is easy to understand in the light of the model we have presented above. The nucleus consists of very strongly interacting neutrons and protons. When the standing wave pattern of these is changed, the nuclear energy stored in the system is changed, with the energy difference being transferred

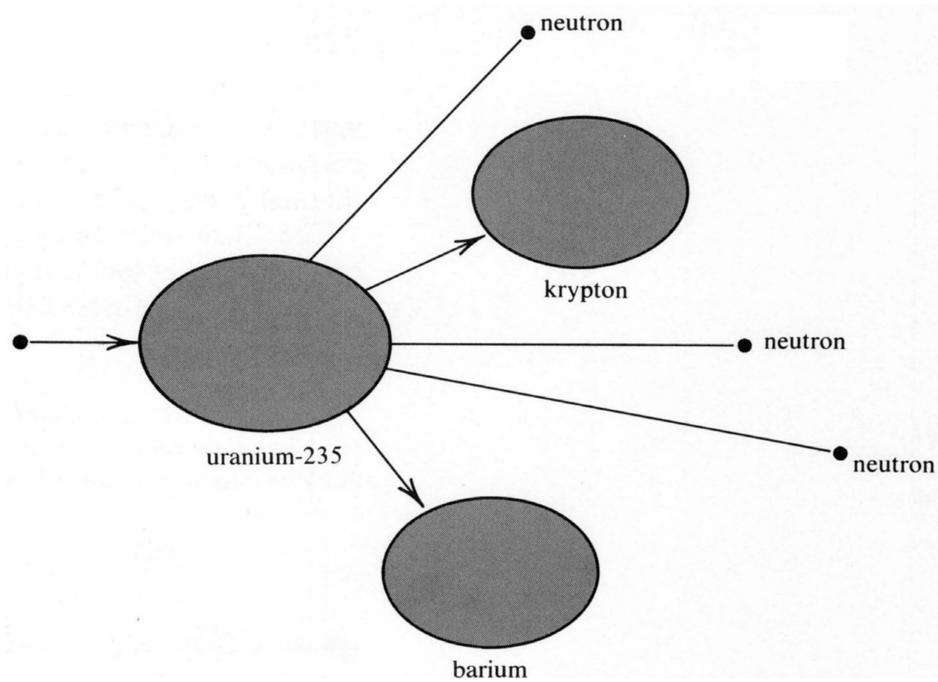


Figure 8.22 Diagram of the nuclear fission process. The isotope uranium-235 is especially susceptible to fission after being struck by a neutron.

to other forms, such as kinetic energy of the nuclear reaction products.

Interestingly enough, the change in energy of a nucleus (as determined by the energy transfer to other forms) also manifests itself as a change in the mass of the nucleus. The energy transfer and the mass change are directly proportional (Eq. 8-11), as predicted by Einstein in his special theory of relativity in 1905.

Equation 8.11
energy transfer = ΔE
change in mass = ΔM_I
speed of light = c

$$\Delta E = \Delta M_I c^2$$

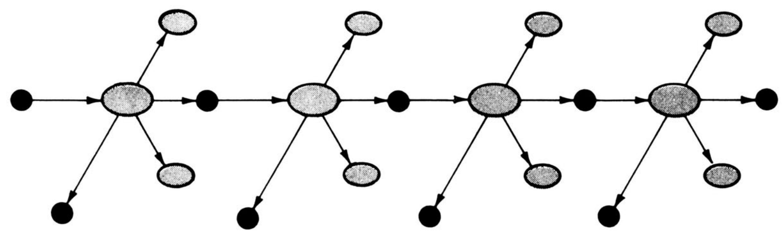
"If a body gives off the energy E in the form of radiation, its mass diminishes by E/c^2. The fact that the energy withdrawn from the body becomes energy of radiation evidently makes no difference, so that we are led to the more general conclusion that the mass of a body is a measure of its energy-content."

Albert Einstein
Annalen der Physik, 1905

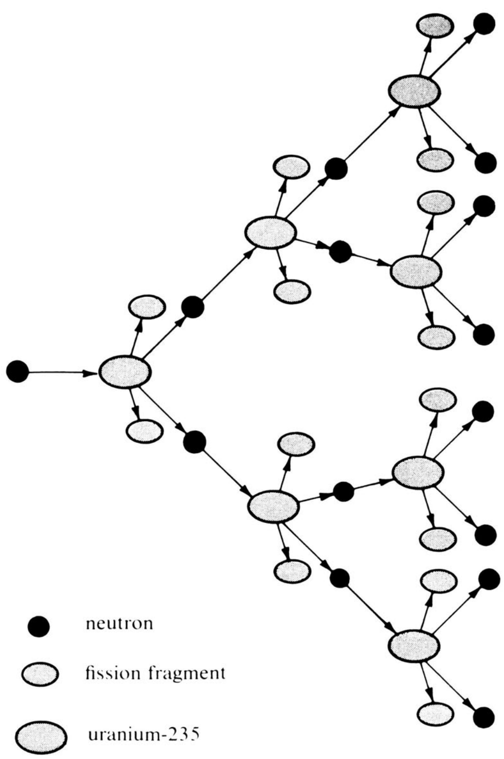
Nuclear fission. *A* nuclear reaction in which an especially large amount of nuclear energy is released is the process of *fission.* Very large nuclei, such as uranium, are capable of fission when they are bombarded by neutrons. In the fission process, a neutron interacts with the uranium nucleus, which then disintegrates into two very energetic nuclear fragments (of about half the mass of uranium) and two or three neutrons (Fig. 8.22). The nuclear fragments transfer their kinetic energy to atoms with which they collide and thereby increase the temperature of the material in which the fissioning nucleus was embedded. In this process, nuclear energy is transformed into thermal energy, with the fission fragments acting as intermediate energy receivers and sources.

Nuclear chain reactions. The utilization of nuclear fission for macroscopic energy transfer requires the fissioning of large numbers of uranium nuclei. This has been accomplished successfully by using the principle of the *chain reaction.* One neutron is required to trigger the fission process, and two or three neutrons are produced. If one of the latter neutrons is allowed to fission a second uranium nucleus, and one of those produced then fissions a third, and so on, the process continues with the result that much nuclear energy is converted to thermal energy (Fig. 8.23a).

The chain reaction can lead to a nuclear explosion, as in an atomic bomb, if two of the neutrons produced during a fission process trigger another fission process (Fig. 8.23b). Then one fission is followed by



(a)



(b)

Figure 8.23 Self-sustaining chain reaction. The neutrons are represented by the small black dots. After a neutron strikes a uranium nucleus, the nucleus is likely to split (fission) into two smaller nuclei and emit other neutrons. These neutrons can be lost (leave the sample or be absorbed by impurities), or they can strike another uranium nucleus.

(a) Non-explosive chain reaction, as in a nuclear power reactor. If some neutrons are lost, so that exactly one neutron from each fission causes another fission, the number of chain reactions is maintained, and the amount of energy released stays constant.

(b) Explosive chain reaction, as in a bomb. If the loss of neutrons is minimized, so that more than one neutron from each fission causes another fission, the number of fission reactions multiplies, and the energy released grows rapidly without stopping. In this diagram, two neutrons from each fission are shown, each of which causes another fission reaction. Thus the number of reactions doubles in each "generation." To see how extraordinarily fast such growth can be, calculate how much money you would have after a month if you earned one penny on the first day, two pennies on the second day, and so on.

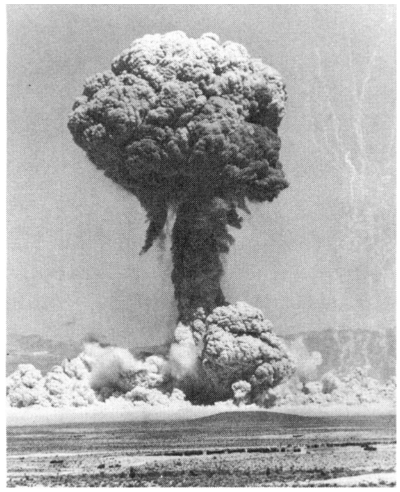


Figure 8.24 Nuclear explosion, Nevada, 1955.

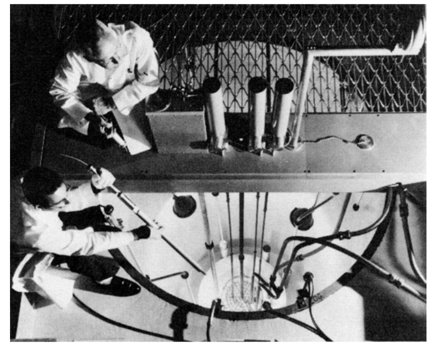


Figure 8.25 In this nuclear reactor for laboratory research, the uranium-containing reactor core is at the bottom of a 24-foot deep tank filled with water. Control rods and measuring devices extend down from the top.

two, the two by four, the four by eight, and so on, until (after about 70 steps) all remaining nuclei fission at once with an enormous energy release equivalent to thousands of tons of TNT (Fig. 8.24).

Nuclear reactors. To harness the chain reaction, it is necessary to have exactly one neutron propagate the process, not more and not less. An excess of neutrons leads to an explosion, a deficiency to extinction of the chain reaction. *A nuclear reactor* is a system in which a steady chain reaction is achieved by careful control of the neutron economy (Fig. 8.25). Some neutrons escape from the system, some neutrons are absorbed by structural members of the reactor, and some neutrons are absorbed without causing fission by cadmium or boron *control rods.* Everything is placed so that exactly one neutron per fission process sustains the reaction. As uranium in the reactor is depleted, the control rods are gradually withdrawn to compensate for the remaining uranium's lowered efficiency. When the reactor is to be shut down, the control rods are inserted more deeply. A control mechanism that carefully monitors the level of neutron production supplies negative feedback via the control rods to counteract any deviations from the desired operating level.

Summary

The many-interacting-particle model for matter had been proposed, discarded, and resurrected several times since the days of the Greek philosophers, when Dalton directed attention at the ratios of weights and volumes in which elements combine to form compounds. Dalton's model was so useful in correlating chemical data that it was accepted after a few decades. Scientists could then turn to the question of the structure of the atoms themselves.

The investigations of electrolysis and spectra gave evidence that matter had an electrical nature and that atoms themselves must be complex systems rather than indivisible entities. The building of models for atoms has revolutionized the physics of the twentieth century. Early attempts involved models made up of particles (electrically negative electrons and the positive nucleus) in electrical interaction with one another. Bohr concluded that the known laws of physics did not apply to atoms, because systems of electrically charged particles could not have the permanence that atoms obviously did have. He therefore assumed new laws, called quantum rules, to supplement the laws of Newton and Maxwell. This approach, while partially successful, soon had to be replaced by a completely new point of view, in which the particle model for matter was abandoned.

Equation 8.1
energy transferred = ΔE
frequency of light = f
Planck's constant = h

$$\Delta E = hf$$

The new concept, introduced by de Broglie and Schrödinger, was to represent the constituents of matter as wave packets instead of as particles. The matter waves are refracted by their interactions with one another, just as particles are deflected by their mutual interactions. Stable atoms are tuned systems in which the electron wave packets oscillate with a characteristic frequency. When the state of the atom changes because radiation is emitted or absorbed, the frequency of the radiation is related to the energy transfer by Bohr's frequency condition (Eq. 8.1).

"The new discoveries made in physics in the last few years, and the ideas and potentialities suggested by them, have had an effect upon the workers in that subject akin to that produced in literature by the Renaissance... In the distance tower still higher peaks, which will yield to those who ascend them still wider prospects, and deepen the feeling, whose truth is emphasized by every advance in science, that `Great are the Works of the Lord."'
J. J. Thomson, 1909

Further developments have led to the recognition that the nucleus of the atom is not indivisible but can undergo spontaneous disintegration in the process called radioactivity. In presently accepted models for the nucleus, protons and neutrons are refracted into a stable standing wave pattern of exceedingly minute physical dimensions by an enormously strong nuclear interaction of very short range. When the state of the nucleus changes, something that happens only during radioactive decay under ordinary conditions on earth, the energy transfer is very much larger than during a change in the electron standing wave pattern in an atom. The technological exploitation of nuclear energy release has led to nuclear reactors for power production as well as to "atomic" (nuclear) bombs.

List of new terms

electrolysis	Planck's constant	gamma rays
electrode	energy level	proton
cathode	ground state	neutron
anode	quantum	atomic number
cathode rays	photoelectric effect	isotope
electron	quantum number	nuclear interaction
electron waves	electron diffraction	nuclear reaction
Thomson's model	wave mechanics	accelerator

alpha rays	valence electron	cyclotron
nucleus	probability	nuclear fission
nuclear model	electron wave packet	chain reaction
planetary model	radioactivity	nuclear reactor
quantum rules	beta rays	control rod

List of symbols

E	energy	$\mathcal{M}$	momentum
ΔE	energy transfer	k	wave number
f	frequency	M_I	inertial mass
h	Planck's constant	v	speed
Δs	size of wave packet (or uncertainty of position)	$\Delta\mathcal{M}$	uncertainty of momentum
		Δk	uncertainty of wave number

Problems

1. Compare the many-interacting-particles models for matter proposed by Greek philosophers, eighteenth-century scientists, and John Dalton. Make reference to: (a) variety in the kinds of particles; (b) interaction among the particles; (c) quantitative relations among the particles.

2. Review and comment upon the evidence that matter has electrically charged constitutents. Which piece of evidence do you find most compelling?

3. Electrons have been described as particles and as waves. Explain briefly what you understand by these terms and how they were appropriate and/or inappropriate.

4. What were some of the problems Bohr was trying to resolve when he introduced his quantum rules? Appraise his success in solving them. (Refer to the Bibliography for additional reading on this subject.)

5. Electron waves are used to study the structure of molecules in a technique called electron diffraction. The inertial mass of electrons is approximately 1 x 10^{-30} kilogram. Find the wavelength of the electron waves in an electron beam when the electron speed is (a) 1.0 x 10^{6} meters per second; (b) 4.5 x 10^{6} meters per second.

6. Would you expect the diffraction of the electron waves in Problem 5 to give evidence of structure on a scale of sizes equal to, larger than, or smaller than visible light? Explain.

7. George Gamow's *Mr. Tompkins Explores the Atom* (published in 1993 as part of *Mr. Tompkins in Paperback*). describes a world where Planck's constant has a numerical value in the macro domain. Read and comment on this science fiction story.

8. Suppose that Planck's constant had the numerical value of 1 kilogram-$(\text{meter})^2$ per second, which is a numerical value in the macro domain. Estimate the consequences of this supposition for the behavior of a few macro-domain objects.

9. Estimate the largest possible value of Planck's constant that is compatible with your everyday experience.

10. Prepare a chronology of events in the development of models for atoms and atomic nuclei between 1900 and 1930.

11. Identify one or more explanations or discussions in this chapter that you find inadequate. Describe the general reasons for your judgment (conclusions contradict your ideas, steps in the reasoning have been omitted, words or phrases are meaningless, equations are hard to follow, . . .), and make your criticism as specific as you can.

Bibliography

I. Adler, *Inside the Nucleus,* Day, New York, 1963.

D. L. Anderson, *The Discovery of the Electron,* Van No strand, Princeton, New Jersey, 1964.

J. Bernstein, *Hitler's Uranium Club, The Secret Recordings at Farm Hall,* American Institute of Physics, Woodbury, New York, 1996.

R. T. Beyer, Ed., *Foundations of Nuclear Physics,* Dover Publications, New York, 1949.

D. Bohm, *Causality and Chance in Modern Physics,* Van Nostrand, Princeton, New Jersey, 1957.

M. Born, *The Restless Universe,* Dover Publications, New York, 1957.

R. Calder, *Living With the Atom,* University of Chicago Press, Chicago, Illinois, 1962.

H. Childs, *An American Genius- The Life of Ernest Orlando Lawrence, E.* P. Dutton, New York, 1968.

C. S. Cook*, Structure of Atomic Nuclei,* Van Nostrand, Princeton, New Jersey, 1964.

E. Curie, *Madame Curie,* Doubleday, Garden City, New York, 1937.

W. C. Dampier and M. Dampier, Editors, *Readings in the Literature of Science,* Harper and Row, New York, 1959.

L. de Broglie, *Matter and Light—The New Physics,* Dover Publications, New York, 1939.

L. de Broglie, *The Revolution in Physics,* Noonday Press, New York, 1953.

A. Eddington, *The Nature of the Physical World,* University of Michigan, Ann Arbor, Michigan, 1958.

Editors of *Scientific American, Atomic Power,* Simon and Schuster, New York, 1956.

A. S. Eve, *Rutherford,* Macmillan, New York, 1939.

L. Fermi, *Atoms in the Family: My Life With Enrico Fermi,* University of Chicago Press, Chicago, 1954.

K. W. Ford, *The World of Elementary Particles,* Blaisdell, Waltham, Massachusetts, 1963.

M. Frayn, *Copenhagen* (play), Methuen, London, 1998.

D. H. Frisch and A. M. Thorndike, *Elementary Particles,* Van Nostrand, Princeton, New Jersey, 1964.

G. Gamow, *Atom and Its Nucleus,* Prentice-Hall, Englewood Cliffs, New Jersey, 1961.

G. Gamow, *Mr. Tompkins Explores the Atom,* Cambridge University Press, New York, 1945.

G. Gamow, *Thirty Years That Shook Physics,* Doubleday, Garden City, New York, 1966.

S. Glasstone, *Source Book on Atomic Energy,* Van Nostrand, Princeton, New Jersey, 1958.

B. Greene, *The Elegant Universe—Superstrings, Hidden Dimensions, and the Quest for the Ultimate Theory,* Vintage Books, New York, 1999.

O. Hahn, *Otto Hahn, A Scientific Autobiography,* Scribner's, New York, 1966.

S. Hecht, *Explaining the Atom,* Viking, New York, 1954.

P. L. Rose, *Heisenberg and the Nazi Atomic Bomb Project, A Study in German Culture*, University of California Press, Berkeley, 1998.

B. Hoffman, *The Strange Story of the Quantum,* Dover Publications, New York, 1959.

J. Irvine, *The Basis of Modern Physics,* Oliver and Boyd, Edinburgh and London, 1968.

G. Johnson, *Strange Beauty—Murray Gell-Mann and the Revolution in Twentieth-Century Physics*, Vintage Books, New York, 1999.

R. Jungk, *Brighter Than a Thousand Suns,* Harcourt Brace, New York, 1958.

W. B. Mann and S. B. Garfinkel, *Radioactivity and Its Measurement,* Van Nostrand, Princeton, New Jersey, 1966.

A. I. Miller, *Einstein, Picasso—Space, Time and the Beauty That Causes Havoc*, Basic Books, New York, 2001.

R. Moore, *Niels Bohr,* Alfred A. Knopf, New York, 1966.

J. Needham and W. Pagel, *Backgrounds to Modern Science,* Macmillan, New York, 1953.

O. Oldenberg, *Introduction to Atomic Physics,* McGraw-Hill, New York, 1954.

H. Smyth, *Atomic Energy for Military Purposes,* Princeton University Press, Princeton, New Jersey, 1945.

A. K. Solomon, *Why Smash Atoms?,* Penguin Books, Harmondsworth, Middlesex, England, 1959.

J. Tyndall, *Faraday as a Discoverer,* Crowell, New York, 1961.

E. Zimmer, *The Revolution in Physics,* Faber and Faber, London, 1936.

Articles from Scientific American. Some or all of these, plus many others, can be obtained on the Internet at http://www.sciamarchive.org/.

H. Alfven, "Antimatter and Cosmology" (April 1967).

L. Badash, "How the `Newer Alchemy' Was Received" (August 1966).

V. D. Barger and D. B. Cline, "High-Energy Scattering" (December 1967).

H. R. Crane, "The *g* Factor of the Electron" (January 1968).

F. J. Dyson, "Mathematics in the Physical Sciences" (September 1964).

G. Feinberg, "Ordinary Matter" (May 1967).

J. F. Hogerton, "The Arrival of Nuclear Power" (February 1968).

R. B. Leachman, "Nuclear Fission" (August 1965).