7. The diagram shows three events and two frames of reference. Describe the time-ordering of the events in these frames. Is any other time-ordering possible in any other frame?

8. The top row shows three clocks located in three different places. They have been synchronized in the frame of reference of the earth, represented by the paper. This synchronization is carried out by exchanging light signals. For example, if the front and back clocks both send out flashes of light when they think it's 2 o'clock, the one in the middle will receive them both at the same time. Event A is the one at which the back clock A reads 2 o'clock, etc. The second row represents clocks that are synchronized aboard the train, which is moving to the right at a substantial fraction of the speed of light. How should the clocks shown with dashed outlines compare with the one at the middle of the train?

9. The figure shows three events and a square representing the t and x axes of a frame of reference. In this frame, the time-ordering of the events is ABC. If we switch to another frame, what other orderings, if any, are possible?

10. The figure shows the motion of an object, with two different frames of reference superimposed. Is anything strange going on?

11. Suppose that the train in 8 is hooked up in a circle, like a chain necklace. What happens?
Exercise 7D: Misconceptions about Relativity

The following is a list of common misconceptions about relativity. The class will be split up into random groups, and each group will cooperate on developing an explanation of the misconception, and then the groups will present their explanations to the class. There may be multiple rounds, with students assigned to different randomly chosen groups in successive rounds.

1. How can light have momentum if it has zero mass?
2. What does the world look like in a frame of reference moving at \( c \)?
3. Alice observes Betty coming toward her from the left at \( c/2 \), and Carol from the right at \( c/2 \). Therefore Betty is moving at the speed of light relative to Carol.
4. Are relativistic effects such as length contraction and time dilation real, or do they just seem to be that way?
5. Special relativity only matters if you’ve moving close to the speed of light.
6. Special relativity says that everything is relative.
7. There is a common misconception that relativistic length contraction is what we would actually see. Refute this by drawing a spacetime diagram for an object approaching an observer, and tracing rays of light emitted from the object’s front and back that both reach the observer’s eye at the same time.
8. When you travel close to the speed of light, your time slows down.
9. Is a light wave’s wavelength relativistically length contracted by a factor of gamma?
10. Accelerate a baseball to ultrarelativistic speeds. Does it become a black hole?
11. Where did the Big Bang happen?
12. The universe can’t be infinite in size, because it’s only had a finite amount of time to expand from the point where the Big Bang happened.
Exercise 7E: The sum of observer-vectors is an observer-vector.

The figure gives four pairs of four-vectors, oriented in our customary way as shown by the light-cone on the left.

1. Of the types shown in the four cases i-iv, which types of vectors could represent the world-line of an observer?

2. Suppose that \( \mathbf{U} \) and \( \mathbf{V} \) are both observer-vectors. What would it mean physically to compute \( \mathbf{U} + \mathbf{V} \)?

3. Determine the sign of each inner product \( \mathbf{A} \cdot \mathbf{B} \).

4. Given an observer whose world-line is along a four-vector \( \mathbf{O} \), suppose we want to determine whether some other four-vector \( \mathbf{P} \) is also a possible world-line of an observer. Show that knowledge of the signs of the inner products \( \mathbf{O} \cdot \mathbf{P} \) and \( \mathbf{P} \cdot \mathbf{P} \) is necessary and sufficient to determine this. Hint: Consider various possibilities like i-iv for vector \( \mathbf{P} \), and see how the signs would turn out.

5. For vectors as described in 4, determine the signs of

\[
(\mathbf{U} + \mathbf{V}) \cdot (\mathbf{U} + \mathbf{V})
\]

and

\[
(\mathbf{U} + \mathbf{V}) \cdot \mathbf{U}
\]

by multiplying them out. Interpret the result physically.
Chapter 8
Atoms and Electromagnetism

8.1 The Electric Glue

Where the telescope ends, the microscope begins. Which of the two has the grander view?

Victor Hugo

His father died during his mother’s pregnancy. Rejected by her as a boy, he was packed off to boarding school when she remarried. He himself never married, but in middle age he formed an intense relationship with a much younger man, a relationship that he terminated when he underwent a psychotic break. Following his early scientific successes, he spent the rest of his professional life mostly in frustration over his inability to unlock the secrets of alchemy.

The man being described is Isaac Newton, but not the triumphant Newton of the standard textbook hagiography. Why dwell on the sad side of his life? To the modern science educator, Newton’s lifelong obsession with alchemy may seem an embarrassment, a distraction from his main achievement, the creation the modern science of mechanics. To Newton, however, his alchemical researches were naturally related to his investigations of force and motion. What was radical about Newton’s analysis of motion was its universality: it succeeded in describing both the heavens and the earth with the same equations, whereas previously it had been assumed that the
sun, moon, stars, and planets were fundamentally different from earthly objects. But Newton realized that if science was to describe all of nature in a unified way, it was not enough to unite the human scale with the scale of the universe: he would not be satisfied until he fit the microscopic universe into the picture as well.

It should not surprise us that Newton failed. Although he was a firm believer in the existence of atoms, there was no more experimental evidence for their existence than there had been when the ancient Greeks first posited them on purely philosophical grounds. Alchemy labored under a tradition of secrecy and mysticism. Newton had already almost single-handedly transformed the fuzzyheaded field of “natural philosophy” into something we would recognize as the modern science of physics, and it would be unjust to criticize him for failing to change alchemy into modern chemistry as well. The time was not ripe. The microscope was a new invention, and it was cutting-edge science when Newton’s contemporary Hooke discovered that living things were made out of cells.

8.1.1 The quest for the atomic force

Newton was not the first of the age of reason. He was the last of the magicians.  
*John Maynard Keynes*

Newton’s quest

Nevertheless it will be instructive to pick up Newton’s train of thought and see where it leads us with the benefit of modern hindsight. In uniting the human and cosmic scales of existence, he had reimagined both as stages on which the actors were objects (trees and houses, planets and stars) that interacted through attractions and repulsions. He was already convinced that the objects inhabiting the microworld were atoms, so it remained only to determine what kinds of forces they exerted on each other.

His next insight was no less brilliant for his inability to bring it to fruition. He realized that the many human-scale forces — friction, sticky forces, the normal forces that keep objects from occupying the same space, and so on — must all simply be expressions of a more fundamental force acting between atoms. Tape sticks to paper because the atoms in the tape attract the atoms in the paper. My house doesn’t fall to the center of the earth because its atoms repel the atoms of the dirt under it.

Here he got stuck. It was tempting to think that the atomic force was a form of gravity, which he knew to be universal, fundamental, and mathematically simple. Gravity, however, is always attractive, so how could he use it to explain the existence of both attractive and repulsive atomic forces? The gravitational force between objects of ordinary size is also extremely small, which is why we never notice cars and houses attracting us gravitationally. It would be hard to understand how gravity could be responsible for anything
as vigorous as the beating of a heart or the explosion of gunpowder. Newton went on to write a million words of alchemical notes filled with speculation about some other force, perhaps a “divine force” or “vegetative force” that would for example be carried by the sperm to the egg.

Luckily, we now know enough to investigate a different suspect as a candidate for the atomic force: electricity. Electric forces are often observed between objects that have been prepared by rubbing (or other surface interactions), for instance when clothes rub against each other in the dryer. A useful example is shown in figure a/1: stick two pieces of tape on a tabletop, and then put two more pieces on top of them. Lift each pair from the table, and then separate them. The two top pieces will then repel each other, a/2, as will the two bottom pieces. A bottom piece will attract a top piece, however, a/3. Electrical forces like these are similar in certain ways to gravity, the other force that we already know to be fundamental:

- Electrical forces are universal. Although some substances, such as fur, rubber, and plastic, respond more strongly to electrical preparation than others, all matter participates in electrical forces to some degree. There is no such thing as a “nonelectric” substance. Matter is both inherently gravitational and inherently electrical.

- Experiments show that the electrical force, like the gravitational force, is an inverse square force. That is, the electrical force between two spheres is proportional to $1/r^2$, where $r$ is the center-to-center distance between them.

Furthermore, electrical forces make more sense than gravity as candidates for the fundamental force between atoms, because we have observed that they can be either attractive or repulsive.

8.1.2 Charge, electricity and magnetism

Charge

“Charge” is the technical term used to indicate that an object has been prepared so as to participate in electrical forces. This is to be distinguished from the common usage, in which the term is used indiscriminately for anything electrical. For example, although we speak colloquially of “charging” a battery, you may easily verify that a battery has no charge in the technical sense, e.g., it does not exert any electrical force on a piece of tape that has been prepared as described in the previous section.
Two types of charge

We can easily collect reams of data on electrical forces between different substances that have been charged in different ways. We find for example that cat fur prepared by rubbing against rabbit fur will attract glass that has been rubbed on silk. How can we make any sense of all this information? A vast simplification is achieved by noting that there are really only two types of charge. Suppose we pick cat fur rubbed on rabbit fur as a representative of type A, and glass rubbed on silk for type B. We will now find that there is no “type C.” Any object electrified by any method is either A-like, attracting things A attracts and repelling those it repels, or B-like, displaying the same attractions and repulsions as B. The two types, A and B, always display opposite interactions. If A displays an attraction with some charged object, then B is guaranteed to undergo repulsion with it, and vice-versa.

The coulomb

Although there are only two types of charge, each type can come in different amounts. The metric unit of charge is the coulomb (rhymes with “drool on”), defined as follows:

One Coulomb (C) is the amount of charge such that a force of \(9.0 \times 10^9\) N occurs between two pointlike objects with charges of 1 C separated by a distance of 1 m.

The notation for an amount of charge is \(q\). The numerical factor in the definition is historical in origin, and is not worth memorizing. The definition is stated for pointlike, i.e., very small, objects, because otherwise different parts of them would be at different distances from each other.

A model of two types of charged particles

Experiments show that all the methods of rubbing or otherwise charging objects involve two objects, and both of them end up getting charged. If one object acquires a certain amount of one type of charge, then the other ends up with an equal amount of the other type. Various interpretations of this are possible, but the simplest is that the basic building blocks of matter come in two flavors, one with each type of charge. Rubbing objects together results in the transfer of some of these particles from one object to the other. In this model, an object that has not been electrically prepared may actually possesses a great deal of both types of charge, but the amounts are equal and they are distributed in the same way throughout it. Since type A repels anything that type B attracts, and vice versa, the object will make a total force of zero on any other object. The rest of this chapter fleshes out this model and discusses how these mysterious particles can be understood as being internal parts of atoms.
Use of positive and negative signs for charge

Because the two types of charge tend to cancel out each other’s forces, it makes sense to label them using positive and negative signs, and to discuss the total charge of an object. It is entirely arbitrary which type of charge to call negative and which to call positive. Benjamin Franklin decided to describe the one we’ve been calling “A” as negative, but it really doesn’t matter as long as everyone is consistent with everyone else. An object with a total charge of zero (equal amounts of both types) is referred to as electrically neutral.

Self-check A

Criticize the following statement: “There are two types of charge, attractive and repulsive.”

Answer, p. 929

A large body of experimental observations can be summarized as follows:

Coulomb’s law: The magnitude of the force acting between point-like charged objects at a center-to-center distance $r$ is given by the equation

$$|F| = k \frac{|q_1||q_2|}{r^2},$$

where the constant $k$ equals $9.0 \times 10^9 \text{ N} \cdot \text{m}^2/\text{C}^2$. The force is attractive if the charges are of different signs, and repulsive if they have the same sign.

Clever modern techniques have allowed the $1/r^2$ form of Coulomb’s law to be tested to incredible accuracy, showing that the exponent is in the range from 1.9999999999999998 to 2.0000000000000002.

Note that Coulomb’s law is closely analogous to Newton’s law of gravity, where the magnitude of the force is $Gm_1m_2/r^2$, except that there is only one type of mass, not two, and gravitational forces are never repulsive. Because of this close analogy between the two types of forces, we can recycle a great deal of our knowledge of gravitational forces. For instance, there is an electrical equivalent of the shell theorem: the electrical forces exerted externally by a uniformly charged spherical shell are the same as if all the charge was concentrated at its center, and the forces exerted internally are zero.

Conservation of indexcharge/conservation of charge

An even more fundamental reason for using positive and negative signs for electrical charge is that experiments show that charge is conserved according to this definition: in any closed system, the total amount of charge is a constant. This is why we observe that rubbing initially uncharged substances together always has the result that one gains a certain amount of one type of charge, while
the other acquires an equal amount of the other type. Conservation of charge seems natural in our model in which matter is made of positive and negative particles. If the charge on each particle is a fixed property of that type of particle, and if the particles themselves can be neither created nor destroyed, then conservation of charge is inevitable.

**Electrical forces involving neutral objects**

As shown in figure b, an electrically charged object can attract objects that are uncharged. How is this possible? The key is that even though each piece of paper has a total charge of zero, it has at least some charged particles in it that have some freedom to move. Suppose that the tape is positively charged, c. Mobile particles in the paper will respond to the tape’s forces, causing one end of the paper to become negatively charged and the other to become positive. The attraction between the paper and the tape is now stronger than the repulsion, because the negatively charged end is closer to the tape.

**self-check B**

What would have happened if the tape was negatively charged?  
Answer, p. 929

**The path ahead**

We have begun to encounter complex electrical behavior that we would never have realized was occurring just from the evidence of our eyes. Unlike the pulleys, blocks, and inclined planes of mechanics, the actors on the stage of electricity and magnetism are invisible phenomena alien to our everyday experience. For this reason, the flavor of the second half of your physics education is dramatically different, focusing much more on experiments and techniques. Even though you will never actually see charge moving through a wire, you can learn to use an ammeter to measure the flow.

Students also tend to get the impression from their first semester of physics that it is a dead science. Not so! We are about to pick up the historical trail that leads directly to the cutting-edge physics research you read about in the newspaper. The atom-smashing experiments that began around 1900, which we will be studying in this chapter, were not that different from the ones of the year 2000 — just smaller, simpler, and much cheaper.

**Magnetic forces**

A detailed mathematical treatment of magnetism won’t come until much later in this book, but we need to develop a few simple ideas about magnetism now because magnetic forces are used in the experiments and techniques we come to next. Everyday magnets
come in two general types. Permanent magnets, such as the ones on your refrigerator, are made of iron or substances like steel that contain iron atoms. (Certain other substances also work, but iron is the cheapest and most common.) The other type of magnet, an example of which is the ones that make your stereo speakers vibrate, consist of coils of wire through which electric charge flows. Both types of magnets are able to attract iron that has not been magnetically prepared, for instance the door of the refrigerator.

A single insight makes these apparently complex phenomena much simpler to understand: magnetic forces are interactions between moving charges, occurring in addition to the electric forces. Suppose a permanent magnet is brought near a magnet of the coiled-wire type. The coiled wire has moving charges in it because we force charge to flow. The permanent magnet also has moving charges in it, but in this case the charges that naturally swirl around inside the iron. (What makes a magnetized piece of iron different from a block of wood is that the motion of the charge in the wood is random rather than organized.) The moving charges in the coiled-wire magnet exert a force on the moving charges in the permanent magnet, and vice-versa.

The mathematics of magnetism is significantly more complex than the Coulomb force law for electricity, which is why we will wait until chapter 11 before delving deeply into it. Two simple facts will suffice for now:

(1) If a charged particle is moving in a region of space near where other charged particles are also moving, their magnetic force on it is directly proportional to its velocity.

(2) The magnetic force on a moving charged particle is always perpendicular to the direction the particle is moving.

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1. **A magnetic compass** example 1

The Earth is molten inside, and like a pot of boiling water, it roils and churns. To make a drastic oversimplification, electric charge can get carried along with the churning motion, so the Earth contains moving charge. The needle of a magnetic compass is itself a small permanent magnet. The moving charge inside the earth interacts magnetically with the moving charge inside the compass needle, causing the compass needle to twist around and point north.

2. **A television tube** example 2

A TV picture is painted by a stream of electrons coming from the back of the tube to the front. The beam scans across the whole surface of the tube like a reader scanning a page of a book. Magnetic forces are used to steer the beam. As the beam comes from the back of the tube to the front, up-down and left-right forces are needed for steering. But magnetic forces cannot be used
to get the beam up to speed in the first place, since they can only push perpendicular to the electrons’ direction of motion, not forward along it.

Discussion Questions

A  If the electrical attraction between two pointlike objects at a distance of 1 m is $9 \times 10^9$ N, why can’t we infer that their charges are $+1$ and $-1$ C? What further observations would we need to do in order to prove this?

B  An electrically charged piece of tape will be attracted to your hand. Does that allow us to tell whether the mobile charged particles in your hand are positive or negative, or both?

8.1.3 Atoms

I was brought up to look at the atom as a nice, hard fellow, red or grey in color according to taste.  

Rutherford

Indexatomismatomism

The Greeks have been kicked around a lot in the last couple of millennia: dominated by the Romans, bullied during the crusades by warlords going to and from the Holy Land, and occupied by Turkey until recently. It’s no wonder they prefer to remember their salad days, when their best thinkers came up with concepts like democracy and atoms. Greece is democratic again after a period of military dictatorship, and an atom is proudly pictured on one of their coins. That’s why it hurts me to have to say that the ancient Greek hypothesis that matter is made of atoms was pure guesswork. There was no real experimental evidence for atoms, and the 18th-century revival of the atom concept by Dalton owed little to the Greeks other than the name, which means “unsplittable.” Subtracting even more cruelly from Greek glory, the name was shown to be inappropriate in 1897 when physicist J.J. Thomson proved experimentally that atoms had even smaller things inside them, which could be extracted. (Thomson called them “electrons.”) The “unsplittable” was splittable after all.

But that’s getting ahead of our story. What happened to the atom concept in the intervening two thousand years? Educated people continued to discuss the idea, and those who were in favor of it could often use it to give plausible explanations for various facts and phenomena. One fact that was readily explained was conservation of mass. For example, if you mix 1 kg of water with 1 kg of dirt, you get exactly 2 kg of mud, no more and no less. The same is true for a variety of processes such as freezing of water, fermenting beer, or pulverizing sandstone. If you believed in atoms, conservation of mass made perfect sense, because all these processes could be interpreted as mixing and rearranging atoms, without changing the total number of atoms. Still, this is nothing like a proof that atoms exist.

If atoms did exist, what types of atoms were there, and what dis-
tinguished the different types from each other? Was it their sizes, their shapes, their weights, or some other quality? The chasm between the ancient and modern atomisms becomes evident when we consider the wild speculations that existed on these issues until the present century. The ancients decided that there were four types of atoms, earth, water, air and fire; the most popular view was that they were distinguished by their shapes. Water atoms were spherical, hence water’s ability to flow smoothly. Fire atoms had sharp points, which was why fire hurt when it touched one’s skin. (There was no concept of temperature until thousands of years later.) The drastically different modern understanding of the structure of atoms was achieved in the course of the revolutionary decade stretching 1895 to 1905. The main purpose of this chapter is to describe those momentous experiments.

Atoms, light, and everything else

Although I tend to ridicule ancient Greek philosophers like Aristotle, let’s take a moment to praise him for something. If you read Aristotle’s writings on physics (or just skim them, which is all I’ve done), the most striking thing is how careful he is about classifying phenomena and analyzing relationships among phenomena. The human brain seems to naturally make a distinction between two types of physical phenomena: objects and motion of objects. When a phenomenon occurs that does not immediately present itself as one of these, there is a strong tendency to conceptualize it as one or the other, or even to ignore its existence completely. For instance, physics teachers shudder at students’ statements that “the dynamite exploded, and force came out of it in all directions.” In these examples, the nonmaterial concept of force is being mentally categorized as if it was a physical substance. The statement that “winding the clock stores motion in the spring” is a miscategorization of electrical energy as a form of motion. An example of ignoring the existence of a phenomenon altogether can be elicited by asking people why we need lamps. The typical response that “the lamp illuminates the room so we can see things,” ignores the necessary role of light coming into our eyes from the things being illuminated.

If you ask someone to tell you briefly about atoms, the likely response is that “everything is made of atoms,” but we’ve now seen that it’s far from obvious which “everything” this statement would properly refer to. For the scientists of the early 1900s who were trying to investigate atoms, this was not a trivial issue of definitions. There was a new gizmo called the vacuum tube, of which the only familiar example today is the picture tube of a TV. In short order, electrical tinkerers had discovered a whole flock of new phenomena that occurred in and around vacuum tubes, and given them picturesque names like “x-rays,” “cathode rays,” “Hertzian waves,” and “N-rays.” These were the types of observations that ended up
telling us that we know about matter, but fierce controversies ensued over whether these were themselves forms of matter.

Let’s bring ourselves up to the level of classification of phenomena employed by physicists in the year 1900. They recognized three categories:

- **Matter** has mass, can have kinetic energy, and can travel through a vacuum, transporting its mass and kinetic energy with it. Matter is conserved, both in the sense of conservation of mass and conservation of the number of atoms of each element. Atoms can’t occupy the same space as other atoms, so a convenient way to prove something is not a form of matter is to show that it can pass through a solid material, in which the atoms are packed together closely.

- **Light** has no mass, always has energy, and can travel through a vacuum, transporting its energy with it. Two light beams can penetrate through each other and emerge from the collision without being weakened, deflected, or affected in any other way. Light can penetrate certain kinds of matter, e.g., glass.

- The third category is everything that doesn’t fit the definition of light or matter. This catch-all category includes, for example, time, velocity, heat, and force.

The chemical index elements, chemical elements

How would one find out what types of atoms there were? Today, it doesn’t seem like it should have been very difficult to work out an experimental program to classify the types of atoms. For each type of atom, there should be a corresponding element, i.e., a pure substance made out of nothing but that type of atom. Atoms are supposed to be unsplittable, so a substance like milk could not possibly be elemental, since churning it vigorously causes it to split up into two separate substances: butter and whey. Similarly, rust could not be an element, because it can be made by combining two substances: iron and oxygen. Despite its apparent reasonableness, no such program was carried out until the eighteenth century. The ancients presumably did not do it because observation was not universally agreed on as the right way to answer questions about nature, and also because they lacked the necessary techniques or the techniques were the province of laborers with low social status, such as smiths and miners. Alchemists were hindered by atomism’s reputation for subversiveness, and by a tendency toward mysticism and secrecy. (The most celebrated challenge facing the alchemists, that of converting lead into gold, is one we now know to be impossible, since lead and gold are both elements.)

By 1900, however, chemists had done a reasonably good job of finding out what the elements were. They also had determined the
ratios of the different atoms’ masses fairly accurately. A typical technique would be to measure how many grams of sodium (Na) would combine with one gram of chlorine (Cl) to make salt (NaCl). (This assumes you’ve already decided based on other evidence that salt consisted of equal numbers of Na and Cl atoms.) The masses of individual atoms, as opposed to the mass ratios, were known only to within a few orders of magnitude based on indirect evidence, and plenty of physicists and chemists denied that individual atoms were anything more than convenient symbols.

Making sense of the elements

As the information accumulated, the challenge was to find a way of systematizing it; the modern scientist’s aesthetic sense rebels against complication. This hodgepodge of elements was an embarrassment. One contemporary observer, William Crookes, described the elements as extending “before us as stretched the wide Atlantic before the gaze of Columbus, mocking, taunting and murmuring strange riddles, which no man has yet been able to solve.” It wasn’t long before people started recognizing that many atoms’ masses were nearly integer multiples of the mass of hydrogen, the lightest element. A few excitable types began speculating that hydrogen was the basic building block, and that the heavier elements were made of clusters of hydrogen. It wasn’t long, however, before their parade was rained on by more accurate measurements, which showed that not all of the elements had atomic masses that were near integer multiples of hydrogen, and even the ones that were close to being integer multiples were off by one percent or so.

Chemistry professor Dmitri Mendeleev, preparing his lectures in 1869, wanted to find some way to organize his knowledge for his students to make it more understandable. He wrote the names of all the elements on cards and began arranging them in different ways on his desk, trying to find an arrangement that would make sense of
the muddle. The row-and-column scheme he came up with is essentially our modern periodic table. The columns of the modern version represent groups of elements with similar chemical properties, and each row is more massive than the one above it. Going across each row, this almost always resulted in placing the atoms in sequence by weight as well. What made the system significant was its predictive value. There were three places where Mendeleev had to leave gaps in his checkerboard to keep chemically similar elements in the same column. He predicted that elements would exist to fill these gaps, and extrapolated or interpolated from other elements in the same column to predict their numerical properties, such as masses, boiling points, and densities. Mendeleev’s professional stock skyrocketed when his three elements (later named gallium, scandium and germanium) were discovered and found to have very nearly the properties he had predicted.

One thing that Mendeleev’s table made clear was that mass was not the basic property that distinguished atoms of different elements. To make his table work, he had to deviate from ordering the elements strictly by mass. For instance, iodine atoms are lighter than tellurium, but Mendeleev had to put iodine after tellurium so that it would lie in a column with chemically similar elements.

**Direct proof that atoms existed**

The success of the kinetic theory of heat was taken as strong evidence that, in addition to the motion of any object as a whole, there is an invisible type of motion all around us: the random motion of atoms within each object. But many conservatives were not convinced that atoms really existed. Nobody had ever seen one, after all. It wasn’t until generations after the kinetic theory of heat was developed that it was demonstrated conclusively that atoms really existed and that they participated in continuous motion that never died out.

The smoking gun to prove atoms were more than mathematical abstractions came when some old, obscure observations were reexamined by an unknown Swiss patent clerk named Albert Einstein. A botanist named Brown, using a microscope that was state of the art in 1827, observed tiny grains of pollen in a drop of water on a microscope slide, and found that they jumped around randomly for no apparent reason. Wondering at first if the pollen he’d assumed to be dead was actually alive, he tried looking at particles of soot, and found that the soot particles also moved around. The same results would occur with any small grain or particle suspended in a liquid. The phenomenon came to be referred to as Brownian motion, and its existence was filed away as a quaint and thoroughly unimportant fact, really just a nuisance for the microscopist.

It wasn’t until 1906 that Einstein found the correct interpreta-
tion for Brown’s observation: the water molecules were in continuous random motion, and were colliding with the particle all the time, kicking it in random directions. After all the millennia of speculation about atoms, at last there was solid proof. Einstein’s calculations dispelled all doubt, since he was able to make accurate predictions of things like the average distance traveled by the particle in a certain amount of time. (Einstein received the Nobel Prize not for his theory of relativity but for his papers on Brownian motion and the photoelectric effect.)

Discussion Questions

A. How could knowledge of the size of an individual aluminum atom be used to infer an estimate of its mass, or vice versa?

B. How could one test Einstein’s interpretation of Brownian motion by observing it at different temperatures?

8.1.4 Quantization of charge

Proving that atoms actually existed was a big accomplishment, but demonstrating their existence was different from understanding their properties. Note that the Brown-Einstein observations had nothing at all to do with electricity, and yet we know that matter is inherently electrical, and we have been successful in interpreting certain electrical phenomena in terms of mobile positively and negatively charged particles. Are these particles atoms? Parts of atoms? Particles that are entirely separate from atoms? It is perhaps premature to attempt to answer these questions without any conclusive evidence in favor of the charged-particle model of electricity.

Strong support for the charged-particle model came from a 1911 experiment by physicist Robert Millikan at the University of Chicago. Consider a jet of droplets of perfume or some other liquid made by blowing it through a tiny pinhole. The droplets emerging from the pinhole must be smaller than the pinhole, and in fact most of them are even more microscopic than that, since the turbulent flow of air tends to break them up. Millikan reasoned that the droplets would acquire a little bit of electric charge as they rubbed against the channel through which they emerged, and if the charged-particle model of electricity was right, the charge might be split up among so many minuscule liquid drops that a single drop might have a total charge amounting to an excess of only a few charged particles — perhaps an excess of one positive particle on a certain drop, or an excess of two negative ones on another.

Millikan’s ingenious apparatus, g, consisted of two metal plates, which could be electrically charged as needed. He sprayed a cloud of oil droplets into the space between the plates, and selected one drop through a microscope for study. First, with no charge on the plates, he would determine the drop’s mass by letting it fall through the air and measuring its terminal velocity, i.e., the velocity at which
the force of air friction canceled out the force of gravity. The force of air drag on a slowly moving sphere had already been found by experiment to be $br^2$, where $b$ was a constant. Setting the total force equal to zero when the drop is at terminal velocity gives

$$bvr^2 - mg = 0,$$

and setting the known density of oil equal to the drop’s mass divided by its volume gives a second equation,

$$
\rho = \frac{m}{\frac{4}{3} \pi r^3}.
$$

Everything in these equations can be measured directly except for $m$ and $r$, so these are two equations in two unknowns, which can be solved in order to determine how big the drop is.

Next Millikan charged the metal plates, adjusting the amount of charge so as to exactly counteract gravity and levitate the drop. If, for instance, the drop being examined happened to have a total charge that was negative, then positive charge put on the top plate would attract it, pulling it up, and negative charge on the bottom plate would repel it, pushing it up. (Theoretically only one plate would be necessary, but in practice a two-plate arrangement like this gave electrical forces that were more uniform in strength throughout the space where the oil drops were.) The amount of charge on the plates required to levitate the charged drop gave Millikan a handle on the amount of charge the drop carried. The more charge the drop had, the stronger the electrical forces on it would be, and the less charge would have to be put on the plates to do the trick. Unfortunately, expressing this relationship using Coulomb’s law would have been impractical, because it would require a perfect knowledge of how the charge was distributed on each plate, plus the ability to perform vector addition of all the forces being exerted on the drop by all the charges on the plate. Instead, Millikan made use of the fact that the electrical force experienced by a pointlike charged object at a certain point in space is proportional to its charge,

$$\frac{F}{q} = \text{constant}.$$  

With a given amount of charge on the plates, this constant could be determined for instance by discarding the oil drop, inserting between the plates a larger and more easily handled object with a known charge on it, and measuring the force with conventional methods. (Millikan actually used a slightly different set of techniques for determining the constant, but the concept is the same.) The amount of force on the actual oil drop had to equal $mg$, since it was just enough to levitate it, and once the calibration constant had been determined, the charge of the drop could then be found based on its previously determined mass.

<table>
<thead>
<tr>
<th>$q$ (C)</th>
<th>$\frac{q}{(1.64 \times 10^{-19} \text{ C})}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$-1.970 \times 10^{-18}$</td>
<td>$-12.02$</td>
</tr>
<tr>
<td>$-0.987 \times 10^{-18}$</td>
<td>$-6.02$</td>
</tr>
<tr>
<td>$-2.773 \times 10^{-18}$</td>
<td>$-16.93$</td>
</tr>
</tbody>
</table>

*H. A few samples of Millikan’s data.*
The table on the left shows a few of the results from Millikan’s 1911 paper. (Millikan took data on both negatively and positively charged drops, but in his paper he gave only a sample of his data on negatively charged drops, so these numbers are all negative.) Even a quick look at the data leads to the suspicion that the charges are not simply a series of random numbers. For instance, the second charge is almost exactly equal to half the first one. Millikan explained the observed charges as all being integer multiples of a single number, $1.64 \times 10^{-19}$ C. In the second column, dividing by this constant gives numbers that are essentially integers, allowing for the random errors present in the experiment. Millikan states in his paper that these results were a 

...direct and tangible demonstration... of the correctness of the view advanced many years ago and supported by evidence from many sources that all electrical charges, however produced, are exact multiples of one definite, elementary electrical charge, or in other words, that an electrical charge instead of being spread uniformly over the charged surface has a definite granular structure, consisting, in fact, of... specks, or atoms of electricity, all precisely alike, peppered over the surface of the charged body.

In other words, he had provided direct evidence for the charged-particle model of electricity and against models in which electricity was described as some sort of fluid. The basic charge is notated $e$, and the modern value is $e = 1.60 \times 10^{-19}$ C. The word “quantized” is used in physics to describe a quantity that can only have certain numerical values, and cannot have any of the values between those. In this language, we would say that Millikan discovered that charge is quantized. The charge $e$ is referred to as the quantum of charge.

A historical note on Millikan’s fraud

Very few undergraduate physics textbooks mention the well-documented fact that although Millikan’s conclusions were correct, he was guilty of scientific fraud. His technique was difficult and painstaking to perform, and his original notebooks, which have been preserved, show that the data were far less perfect than he claimed in his published scientific papers. In his publications, he stated categorically that every single oil drop observed had had a charge that was a multiple of $e$, with no exceptions or omissions. But his notebooks are replete with notations such as “beautiful data, keep,” and “bad run, throw out.” Millikan, then, appears to have earned his Nobel Prize by advocating a correct position with dishonest descriptions of his data.

Why do textbook authors fail to mention Millikan’s fraud? It may be that they think students are too unsophisticated to cor-
directly evaluate the implications of the fact that scientific fraud has sometimes existed and even been rewarded by the scientific establishment. Maybe they are afraid students will reason that fudging data is OK, since Millikan got the Nobel Prize for it. But falsifying history in the name of encouraging truthfulness is more than a little ironic. English teachers don’t edit Shakespeare’s tragedies so that the bad characters are always punished and the good ones never suffer!

*self-check C*

Is money quantized? What is the quantum of money?  
▷ Answer, p. 929

8.1.5 The electron

Cathode rays

Nineteenth-century physicists spent a lot of time trying to come up with wild, random ways to play with electricity. The best experiments of this kind were the ones that made big sparks or pretty colors of light.

One such parlor trick was the cathode ray. To produce it, you first had to hire a good glassblower and find a good vacuum pump. The glassblower would create a hollow tube and embed two pieces of metal in it, called the electrodes, which were connected to the outside via metal wires passing through the glass. Before letting him seal up the whole tube, you would hook it up to a vacuum pump, and spend several hours huffing and puffing away at the pump’s hand crank to get a good vacuum inside. Then, while you were still pumping on the tube, the glassblower would melt the glass and seal the whole thing shut. Finally, you would put a large amount of positive charge on one wire and a large amount of negative charge on the other. Metals have the property of letting charge move through them easily, so the charge deposited on one of the wires would quickly spread out because of the repulsion of each part of it for every other part. This spreading-out process would result in nearly all the charge ending up in the electrodes, where there is more room to spread out than there is in the wire. For obscure historical reasons a negative electrode is called a cathode and a positive one is an anode.

Figure i shows the light-emitting stream that was observed. If, as shown in this figure, a hole was made in the anode, the beam would extend on through the hole until it hit the glass. Drilling a hole in the cathode, however would not result in any beam coming out on the left side, and this indicated that the stuff, whatever it was, was coming from the cathode. The rays were therefore christened “cathode rays.” (The terminology is still used today in the term “cathode ray tube” or “CRT” for the picture tube of a TV or computer monitor.)
Were cathode rays a form of light, or of matter?

Were cathode rays a form of light, or matter? At first no one really cared what they were, but as their scientific importance became more apparent, the light-versus-matter issue turned into a controversy along nationalist lines, with the Germans advocating light and the English holding out for matter. The supporters of the material interpretation imagined the rays as consisting of a stream of atoms ripped from the substance of the cathode.

One of our defining characteristics of matter is that material objects cannot pass through each other. Experiments showed that cathode rays could penetrate at least some small thickness of matter, such as a metal foil a tenth of a millimeter thick, implying that they were a form of light.

Other experiments, however, pointed to the contrary conclusion. Light is a wave phenomenon, and one distinguishing property of waves is demonstrated by speaking into one end of a paper towel roll. The sound waves do not emerge from the other end of the tube as a focused beam. Instead, they begin spreading out in all directions as soon as they emerge. This shows that waves do not necessarily travel in straight lines. If a piece of metal foil in the shape of a star or a cross was placed in the way of the cathode ray, then a “shadow” of the same shape would appear on the glass, showing that the rays traveled in straight lines. This straight-line motion suggested that they were a stream of small particles of matter.

These observations were inconclusive, so what was really needed was a determination of whether the rays had mass and weight. The trouble was that cathode rays could not simply be collected in a cup and put on a scale. When the cathode ray tube is in operation, one does not observe any loss of material from the cathode, or any crust being deposited on the anode.

Nobody could think of a good way to weigh cathode rays, so the next most obvious way of settling the light/matter debate was to check whether the cathode rays possessed electrical charge. Light was known to be uncharged. If the cathode rays carried charge, they were definitely matter and not light, and they were presumably being made to jump the gap by the simultaneous repulsion of the negative charge in the cathode and attraction of the positive charge in the anode. The rays would overshoot the anode because of their momentum. (Although electrically charged particles do not normally leap across a gap of vacuum, very large amounts of charge were being used, so the forces were unusually intense.)

Thomson’s experiments

Physicist J.J. Thomson at Cambridge carried out a series of definitive experiments on cathode rays around the year 1897. By turning them slightly off course with electrical forces, k, he showed...
that they were indeed electrically charged, which was strong evidence that they were material. Not only that, but he proved that they had mass, and measured the ratio of their mass to their charge, \( m/q \). Since their mass was not zero, he concluded that they were a form of matter, and presumably made up of a stream of microscopic, negatively charged particles. When Millikan published his results fourteen years later, it was reasonable to assume that the charge of one such particle equaled minus one fundamental charge, \( q = -e \), and from the combination of Thomson’s and Millikan’s results one could therefore determine the mass of a single cathode ray particle.

The basic technique for determining \( m/q \) was simply to measure the angle through which the charged plates bent the beam. The electric force acting on a cathode ray particle while it was between the plates would be proportional to its charge,

\[
F_{\text{elec}} = (\text{known constant}) \cdot q
\]

Application of Newton’s second law, \( a = F/m \), would allow \( m/q \) to be determined:

\[
\frac{m}{q} = \frac{\text{known constant}}{a}
\]

There was just one catch. Thomson needed to know the cathode ray particles’ velocity in order to figure out their acceleration. At that point, however, nobody had even an educated guess as to the speed of the cathode rays produced in a given vacuum tube. The beam appeared to leap across the vacuum tube practically instantaneously, so it was no simple matter of timing it with a stopwatch!

Thomson’s clever solution was to observe the effect of both electric and magnetic forces on the beam. The magnetic force exerted by a particular magnet would depend on both the cathode ray’s charge and its velocity:

\[
F_{\text{mag}} = (\text{known constant} \ #2) \cdot qv
\]
Thomson played with the electric and magnetic forces until either one would produce an equal effect on the beam, allowing him to solve for the velocity,

\[ v = \frac{\text{(known constant)}}{\text{(known constant #2)}}. \]

Knowing the velocity (which was on the order of 10% of the speed of light for his setup), he was able to find the acceleration and thus the mass-to-charge ratio \( m/q \). Thomson’s techniques were relatively crude (or perhaps more charitably we could say that they stretched the state of the art of the time), so with various methods he came up with \( m/q \) values that ranged over about a factor of two, even for cathode rays extracted from a cathode made of a single material. The best modern value is \( m/q = 5.69 \times 10^{-12} \) kg/C, which is consistent with the low end of Thomson’s range.

The cathode ray as a subatomic particle: the indexelectron electron

What was significant about Thomson’s experiment was not the actual numerical value of \( m/q \), however, so much as the fact that, combined with Millikan’s value of the fundamental charge, it gave a mass for the cathode ray particles that was thousands of times smaller than the mass of even the lightest atoms. Even without Millikan’s results, which were 14 years in the future, Thomson recognized that the cathode rays’ \( m/q \) was thousands of times smaller than the \( m/q \) ratios that had been measured for electrically charged atoms in chemical solutions. He correctly interpreted this as evidence that the cathode rays were smaller building blocks — he called them electrons — out of which atoms themselves were formed. This was an extremely radical claim, coming at a time when atoms had not yet been proven to exist! Even those who used the word “atom” often considered them no more than mathematical abstractions, not literal objects. The idea of searching for structure inside of “unsplittable” atoms was seen by some as lunacy, but within ten years Thomson’s ideas had been amply verified by many more detailed experiments.

Discussion Questions

A Thomson started to become convinced during his experiments that the “cathode rays” observed coming from the cathodes of vacuum tubes were building blocks of atoms — what we now call electrons. He then carried out observations with cathodes made of a variety of metals, and found that \( m/q \) was roughly the same in every case, considering his limited accuracy. Given his suspicion, why did it make sense to try different metals? How would the consistent values of \( m/q \) serve to test his hypothesis?

B My students have frequently asked whether the \( m/q \) that Thomson measured was the value for a single electron, or for the whole beam. Can you answer this question?
Thomson found that the $m/q$ of an electron was thousands of times smaller than that of charged atoms in chemical solutions. Would this imply that the electrons had more charge? Less mass? Would there be no way to tell? Explain. Remember that Millikan’s results were still many years in the future, so $q$ was unknown.

Can you guess any practical reason why Thomson couldn’t just let one electron fly across the gap before disconnecting the battery and turning off the beam, and then measure the amount of charge deposited on the anode, thus allowing him to measure the charge of a single electron directly?

Why is it not possible to determine $m$ and $q$ themselves, rather than just their ratio, by observing electrons’ motion in electric and magnetic fields?

### 8.1.6 The raisin cookie model of the atom

Based on his experiments, Thomson proposed a picture of the atom which became known as the raisin cookie model. In the neutral atom, there are four electrons with a total charge of $-4e$, sitting in a sphere (the “cookie”) with a charge of $+4e$ spread throughout it. It was known that chemical reactions could not change one element into another, so in Thomson’s scenario, each element’s cookie sphere had a permanently fixed radius, mass, and positive charge, different from those of other elements. The electrons, however, were not a permanent feature of the atom, and could be tacked on or pulled out to make charged ions. Although we now know, for instance, that a neutral atom with four electrons is the element beryllium, scientists at the time did not know how many electrons the various neutral atoms possessed.

This model is clearly different from the one you’ve learned in grade school or through popular culture, where the positive charge is concentrated in a tiny nucleus at the atom’s center. An equally important change in ideas about the atom has been the realization that atoms and their constituent subatomic particles behave entirely differently from objects on the human scale. For instance, we’ll see later that an electron can be in more than one place at one time. The raisin cookie model was part of a long tradition of attempts to make mechanical models of phenomena, and Thomson and his contemporaries never questioned the appropriateness of building a mental model of an atom as a machine with little parts inside. Today, mechanical models of atoms are still used (for instance the tinker-toy-style molecular modeling kits like the ones used by Watson and Crick to figure out the double helix structure of DNA), but scientists realize that the physical objects are only aids to help our brains’ symbolic and visual processes think about atoms.

Although there was no clear-cut experimental evidence for many of the details of the raisin cookie model, physicists went ahead and started working out its implications. For instance, suppose you had
a four-electron atom. All four electrons would be repelling each other, but they would also all be attracted toward the center of the “cookie” sphere. The result should be some kind of stable, symmetric arrangement in which all the forces canceled out. People sufficiently clever with math soon showed that the electrons in a four-electron atom should settle down at the vertices of a pyramid with one less side than the Egyptian kind, i.e., a regular tetrahedron. This deduction turns out to be wrong because it was based on incorrect features of the model, but the model also had many successes, a few of which we will now discuss.

Flow of electrical charge in wires example 3
One of my former students was the son of an electrician, and had become an electrician himself. He related to me how his father had remained refused to believe all his life that electrons really flowed through wires. If they had, he reasoned, the metal would have gradually become more and more damaged, eventually crumbling to dust.

His opinion is not at all unreasonable based on the fact that electrons are material particles, and that matter cannot normally pass through matter without making a hole through it. Nineteenth-century physicists would have shared his objection to a charged-particle model of the flow of electrical charge. In the raisin-cookie model, however, the electrons are very low in mass, and therefore presumably very small in size as well. It is not surprising that they can slip between the atoms without damaging them.

Flow of electrical charge across cell membranes example 4
Your nervous system is based on signals carried by charge moving from nerve cell to nerve cell. Your body is essentially all liquid, and atoms in a liquid are mobile. This means that, unlike the case of charge flowing in a solid wire, entire charged atoms can flow in your nervous system.

Emission of electrons in a cathode ray tube example 5
Why do electrons detach themselves from the cathode of a vacuum tube? Certainly they are encouraged to do so by the repulsion of the negative charge placed on the cathode and the attraction from the net positive charge of the anode, but these are not strong enough to rip electrons out of atoms by main force — if they were, then the entire apparatus would have been instantly vaporized as every atom was simultaneously ripped apart!

The raisin cookie model leads to a simple explanation. We know that heat is the energy of random motion of atoms. The atoms in any object are therefore violently jostling each other all the time, and a few of these collisions are violent enough to knock electrons out of atoms. If this occurs near the surface of a solid object, the electron may come loose. Ordinarily, however, this loss of electrons is a self-limiting process; the loss of electrons leaves
the object with a net positive charge, which attracts the lost sheep home to the fold. (For objects immersed in air rather than vacuum, there will also be a balanced exchange of electrons between the air and the object.)

This interpretation explains the warm and friendly yellow glow of the vacuum tubes in an antique radio. To encourage the emission of electrons from the vacuum tubes’ cathodes, the cathodes are intentionally warmed up with little heater coils.

Discussion Questions

A Today many people would define an ion as an atom (or molecule) with missing electrons or extra electrons added on. How would people have defined the word “ion” before the discovery of the electron?

B Since electrically neutral atoms were known to exist, there had to be positively charged subatomic stuff to cancel out the negatively charged electrons in an atom. Based on the state of knowledge immediately after the Millikan and Thomson experiments, was it possible that the positively charged stuff had an unquantized amount of charge? Could it be quantized in units of +e? In units of +2e? In units of +5/7e?

This chapter is summarized on page 959. Notation and terminology are tabulated on pages 945-946.

8.2 The Nucleus

8.2.1 Radioactivity

Becquerel’s discovery of radioactivity

How did physicists figure out that the raisin cookie model was incorrect, and that the atom’s positive charge was concentrated in a tiny, central nucleus? The story begins with the discovery of radioactivity by the French chemist Becquerel. Up until radioactivity was discovered, all the processes of nature were thought to be based on chemical reactions, which were rearrangements of combinations of atoms. Atoms exert forces on each other when they are close together, so sticking or unsticking them would either release or store electrical energy. That energy could be converted to and from other forms, as when a plant uses the energy in sunlight to make sugars and carbohydrates, or when a child eats sugar, releasing the energy in the form of kinetic energy.

Becquerel discovered a process that seemed to release energy from an unknown new source that was not chemical. Becquerel, whose father and grandfather had also been physicists, spent the first twenty years of his professional life as a successful civil engineer, teaching physics on a part-time basis. He was awarded the chair of physics at the Musée d’Histoire Naturelle in Paris after the death of his father, who had previously occupied it. Having now a significant amount of time to devote to physics, he began studying the interaction of light and matter. He became interested in the phe-
nomenon of phosphorescence, in which a substance absorbs energy from light, then releases the energy via a glow that only gradually goes away. One of the substances he investigated was a uranium compound, the salt UKSO$_5$. One day in 1896, cloudy weather interfered with his plan to expose this substance to sunlight in order to observe its fluorescence. He stuck it in a drawer, coincidentally on top of a blank photographic plate — the old-fashioned glass-backed counterpart of the modern plastic roll of film. The plate had been carefully wrapped, but several days later when Becquerel checked it in the darkroom before using it, he found that it was ruined, as if it had been completely exposed to light.

History provides many examples of scientific discoveries that happened this way: an alert and inquisitive mind decides to investigate a phenomenon that most people would not have worried about explaining. Becquerel first determined by further experiments that the effect was produced by the uranium salt, despite a thick wrapping of paper around the plate that blocked out all light. He tried a variety of compounds, and found that it was the uranium that did it: the effect was produced by any uranium compound, but not by any compound that didn’t include uranium atoms. The effect could be at least partially blocked by a sufficient thickness of metal, and he was able to produce silhouettes of coins by interposing them between the uranium and the plate. This indicated that the effect traveled in a straight line, so that it must have been some kind of ray rather than, e.g., the seepage of chemicals through the paper. He used the word “radiations,” since the effect radiated out from the uranium salt.

At this point Becquerel still believed that the uranium atoms were absorbing energy from light and then gradually releasing the energy in the form of the mysterious rays, and this was how he presented it in his first published lecture describing his experiments. Interesting, but not earth-shattering. But he then tried to determine how long it took for the uranium to use up all the energy that had supposedly been stored in it by light, and he found that it never seemed to become inactive, no matter how long he waited. Not only that, but a sample that had been exposed to intense sunlight for a whole afternoon was no more or less effective than a sample that had always been kept inside. Was this a violation of conservation of energy? If the energy didn’t come from exposure to light, where did it come from?

*Three kinds of “radiations”*

Unable to determine the source of the energy directly, turn-of-the-century physicists instead studied the behavior of the “radiations” once they had been emitted. Becquerel had already shown that the radioactivity could penetrate through cloth and paper, so the first obvious thing to do was to investigate in more detail what
thickness of material the radioactivity could get through. They soon learned that a certain fraction of the radioactivity's intensity would be eliminated by even a few inches of air, but the remainder was not eliminated by passing through more air. Apparently, then, the radioactivity was a mixture of more than one type, of which one was blocked by air. They then found that of the part that could penetrate air, a further fraction could be eliminated by a piece of paper or a very thin metal foil. What was left after that, however, was a third, extremely penetrating type, some of whose intensity would still remain even after passing through a brick wall. They decided that this showed there were three types of radioactivity, and without having the faintest idea of what they really were, they made up names for them. The least penetrating type was arbitrarily labeled $\alpha$ (alpha), the first letter of the Greek alphabet, and so on through $\beta$ (beta) and finally $\gamma$ (gamma) for the most penetrating type.

Radium: a more intense source of radioactivity

The measuring devices used to detect radioactivity were crude: photographic plates or even human eyeballs (radioactivity makes flashes of light in the jelly-like fluid inside the eye, which can be seen by the eyeball's owner if it is otherwise very dark). Because the ways of detecting radioactivity were so crude and insensitive, further progress was hindered by the fact that the amount of radioactivity emitted by uranium was not really very great. The vital contribution of physicist/chemist Marie Curie and her husband Pierre was to discover the element radium, and to purify and isolate significant quantities it. Radium emits about a million times more radioactivity per unit mass than uranium, making it possible to do the experiments that were needed to learn the true nature of radioactivity. The dangers of radioactivity to human health were then unknown, and Marie died of leukemia thirty years later. (Pierre was run over and killed by a horsecart.)

Tracking down the nature of alphas, betas, and gammas

As radium was becoming available, an apprentice scientist named Ernest Rutherford arrived in England from his native New Zealand and began studying radioactivity at the Cavendish Laboratory. The young colonial’s first success was to measure the mass-to-charge ratio of beta rays. The technique was essentially the same as the one Thomson had used to measure the mass-to-charge ratio of cathode rays by measuring their deflections in electric and magnetic fields. The only difference was that instead of the cathode of a vacuum tube, a nugget of radium was used to supply the beta rays. Not only was the technique the same, but so was the result. Beta rays had the same $m/q$ ratio as cathode rays, which suggested they were one and the same. Nowadays, it would make sense simply to use the term “electron,” and avoid the archaic “cathode ray” and “beta particle,” but the old labels are still widely used, and it is unfortu-
nately necessary for physics students to memorize all three names for the same thing.

At first, it seemed that neither alphas or gammas could be deflected in electric or magnetic fields, making it appear that neither was electrically charged. But soon Rutherford obtained a much more powerful magnet, and was able to use it to deflect the alphas but not the gammas. The alphas had a much larger value of \( m/q \) than the betas (about 4000 times greater), which was why they had been so hard to deflect. Gammas are uncharged, and were later found to be a form of light.

The \( m/q \) ratio of alpha particles turned out to be the same as those of two different types of ions, \( \text{He}^{++} \) (a helium atom with two missing electrons) and \( \text{H}_2^+ \) (two hydrogen atoms bonded into a molecule, with one electron missing), so it seemed likely that they were one or the other of those. The diagram shows a simplified version of Rutherford’s ingenious experiment proving that they were \( \text{He}^{++} \) ions. The gaseous element radon, an alpha emitter, was introduced into one half of a double glass chamber. The glass wall dividing the chamber was made extremely thin, so that some of the rapidly moving alpha particles were able to penetrate it. The other chamber, which was initially evacuated, gradually began to accumulate a population of alpha particles (which would quickly pick up electrons from their surroundings and become electrically neutral). Rutherford then determined that it was helium gas that had appeared in the second chamber. Thus alpha particles were proved to be \( \text{He}^{++} \) ions. The nucleus was yet to be discovered, but in modern terms, we would describe a \( \text{He}^{++} \) ion as the nucleus of a He atom.

To summarize, here are the three types of radiation emitted by radioactive elements, and their descriptions in modern terms:

<table>
<thead>
<tr>
<th>( \alpha ) particle</th>
<th>stopped by a few inches of air</th>
<th>( \text{He} ) nucleus</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \beta ) particle</td>
<td>stopped by a piece of paper</td>
<td>electron</td>
</tr>
<tr>
<td>( \gamma ) ray</td>
<td>penetrates thick shielding</td>
<td>a type of light</td>
</tr>
</tbody>
</table>

**Discussion Question**

A Most sources of radioactivity emit alphas, betas, and gammas, not just one of the three. In the radon experiment, how did Rutherford know that he was studying the alphas?

8.2.2 The planetary model

The stage was now set for the unexpected discovery that the positively charged part of the atom was a tiny, dense lump at the atom’s center rather than the “cookie dough” of the raisin cookie model. By 1909, Rutherford was an established professor, and had students working under him. For a raw undergraduate named Marsden, he picked a research project he thought would be tedious but straightforward.
It was already known that although alpha particles would be stopped completely by a sheet of paper, they could pass through a sufficiently thin metal foil. Marsden was to work with a gold foil only 1000 atoms thick. (The foil was probably made by evaporating a little gold in a vacuum chamber so that a thin layer would be deposited on a glass microscope slide. The foil would then be lifted off the slide by submerging the slide in water.)

Rutherford had already determined in his previous experiments the speed of the alpha particles emitted by radium, a fantastic $1.5 \times 10^7$ m/s. The experimenters in Rutherford’s group visualized them as very small, very fast cannonballs penetrating the “cookie dough” part of the big gold atoms. A piece of paper has a thickness of a hundred thousand atoms or so, which would be sufficient to stop them completely, but crashing through a thousand would only slow them a little and turn them slightly off of their original paths.

Marsden’s supposedly ho-hum assignment was to use the apparatus shown in figure f to measure how often alpha particles were deflected at various angles. A tiny lump of radium in a box emitted alpha particles, and a thin beam was created by blocking all the alphas except those that happened to pass out through a tube. Typically deflected in the gold by only a small amount, they would reach a screen very much like the screen of a TV’s picture tube, which would make a flash of light when it was hit. Here is the first example we have encountered of an experiment in which a beam of particles is detected one at a time. This was possible because each alpha particle carried so much kinetic energy; they were moving at about the same speed as the electrons in the Thomson experiment, but had ten thousand times more mass.

Marsden sat in a dark room, watching the apparatus hour after hour and recording the number of flashes with the screen moved to various angles. The rate of the flashes was highest when he set the screen at an angle close to the line of the alphas’ original path, but if he watched an area farther off to the side, he would also occasionally see an alpha that had been deflected through a larger angle. After seeing a few of these, he got the crazy idea of moving the screen to see if even larger angles ever occurred, perhaps even angles larger than 90 degrees.

The crazy idea worked: a few alpha particles were deflected through angles of up to 180 degrees, and the routine experiment had become an epoch-making one. Rutherford said, “We have been able to get some of the alpha particles coming backwards. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you.” Explanations were hard to come by in the raisin cookie model. What intense electrical forces
could have caused some of the alpha particles, moving at such astronomical speeds, to change direction so drastically? Since each gold atom was electrically neutral, it would not exert much force on an alpha particle outside it. True, if the alpha particle was very near to or inside of a particular atom, then the forces would not necessarily cancel out perfectly; if the alpha particle happened to come very close to a particular electron, the $1/r^2$ form of the Coulomb force law would make for a very strong force. But Marsden and Rutherford knew that an alpha particle was 8000 times more massive than an electron, and it is simply not possible for a more massive object to rebound backwards from a collision with a less massive object while conserving momentum and energy. It might be possible in principle for a particular alpha to follow a path that took it very close to one electron, and then very close to another electron, and so on, with the net result of a large deflection, but careful calculations showed that such multiple “close encounters” with electrons would be millions of times too rare to explain what was actually observed.

At this point, Rutherford and Marsden dusted off an unpopular and neglected model of the atom, in which all the electrons orbited around a small, positively charged core or “nucleus,” just like the planets orbiting around the sun. All the positive charge and nearly all the mass of the atom would be concentrated in the nucleus, rather than spread throughout the atom as in the raisin cookie model. The positively charged alpha particles would be repelled by the gold atom’s nucleus, but most of the alphas would not come close enough to any nucleus to have their paths drastically altered. The few that did come close to a nucleus, however, could rebound backwards from a single such encounter, since the nucleus of a heavy gold atom would be fifty times more massive than an alpha...
particle. It turned out that it was not even too difficult to derive a formula giving the relative frequency of deflections through various angles, and this calculation agreed with the data well enough (to within 15%), considering the difficulty in getting good experimental statistics on the rare, very large angles.

What had started out as a tedious exercise to get a student started in science had ended as a revolution in our understanding of nature. Indeed, the whole thing may sound a little too much like a moralistic fable of the scientific method with overtones of the Horatio Alger genre. The skeptical reader may wonder why the planetary model was ignored so thoroughly until Marsden and Rutherford’s discovery. Is science really more of a sociological enterprise, in which certain ideas become accepted by the establishment, and other, equally plausible explanations are arbitrarily discarded? Some social scientists are currently ruffling a lot of scientists’ feathers with critiques very much like this, but in this particular case, there were very sound reasons for rejecting the planetary model. As you’ll learn in more detail later in this course, any charged particle that undergoes an acceleration dissipate energy in the form of light. In the planetary model, the electrons were orbiting the nucleus in circles or ellipses, which meant they were undergoing acceleration, just like the acceleration you feel in a car going around a curve. They should have dissipated energy as light, and eventually they should have lost all their energy. Atoms don’t spontaneously collapse like that, which was why the raisin cookie model, with its stationary electrons, was originally preferred. There were other problems as well. In the planetary model, the one-electron atom would have to be flat, which would be inconsistent with the success of molecular modeling with spherical balls representing hydrogen and atoms. These molecular models also seemed to work best if specific sizes were used for different atoms, but there is no obvious reason in the planetary model why the radius of an electron’s orbit should be a fixed number. In view of the conclusive Marsden-Rutherford results, however, these became fresh puzzles in atomic physics, not reasons for disbelieving the planetary model.

Some phenomena explained with the planetary model

The planetary model may not be the ultimate, perfect model of the atom, but don’t underestimate its power. It already allows us to visualize correctly a great many phenomena.

As an example, let’s consider the distinctions among nonmetals, metals that are magnetic, and metals that are nonmagnetic. As shown in figure i, a metal differs from a nonmetal because its outermost electrons are free to wander rather than owing their allegiance to a particular atom. A metal that can be magnetized is one that is willing to line up the rotations of some of its electrons so that their axes are parallel. Recall that magnetic forces are forces made
by moving charges; we have not yet discussed the mathematics and geometry of magnetic forces, but it is easy to see how random orientations of the atoms in the nonmagnetic substance would lead to cancellation of the forces.

Even if the planetary model does not immediately answer such questions as why one element would be a metal and another a nonmetal, these ideas would be difficult or impossible to conceptualize in the raisin cookie model.

Discussion Question

A. In reality, charges of the same type repel one another and charges of different types are attracted. Suppose the rules were the other way around, giving repulsion between opposite charges and attraction between similar ones. What would the universe be like?

8.2.3 Atomic number

As alluded to in a discussion question in the previous section, scientists of this period had only a very approximate idea of how many units of charge resided in the nuclei of the various chemical elements. Although we now associate the number of units of nuclear charge with the element’s position on the periodic table, and call it the atomic number, they had no idea that such a relationship existed. Mendeleev’s table just seemed like an organizational tool, not something with any necessary physical significance. And everything Mendeleev had done seemed equally valid if you turned the table upside-down or reversed its left and right sides, so even if you wanted to number the elements sequentially with integers, there was an ambiguity as to how to do it. Mendeleev’s original table was in fact upside-down compared to the modern one.

In the period immediately following the discovery of the nucleus, physicists only had rough estimates of the charges of the various nuclei. In the case of the very lightest nuclei, they simply found the maximum number of electrons they could strip off by various methods: chemical reactions, electric sparks, ultraviolet light, and so on.
For example they could easily strip off one or two electrons from helium, making He$^+$ or He$^{++}$, but nobody could make He$^{+++}$, presumably because the nuclear charge of helium was only $+2e$. Unfortunately only a few of the lightest elements could be stripped completely, because the more electrons were stripped off, the greater the positive net charge remaining, and the more strongly the rest of the negatively charged electrons would be held on. The heavy elements’ atomic numbers could only be roughly extrapolated from the light elements, where the atomic number was about half the atom’s mass expressed in units of the mass of a hydrogen atom. Gold, for example, had a mass about 197 times that of hydrogen, so its atomic number was estimated to be about half that, or somewhere around 100. We now know it to be 79.

How did we finally find out? The riddle of the nuclear charges was at last successfully attacked using two different techniques, which gave consistent results. One set of experiments, involving x-rays, was performed by the young Henry Mosely, whose scientific brilliance was soon to be sacrificed in a battle between European imperialists over who would own the Dardanelles, during that pointless conflict then known as the War to End All Wars, and now referred to as World War I.

Since Mosely’s analysis requires several concepts with which you are not yet familiar, we will instead describe the technique used by James Chadwick at around the same time. An added bonus of describing Chadwick’s experiments is that they presaged the important modern technique of studying collisions of subatomic particles. In grad school, I worked with a professor whose thesis adviser’s thesis adviser was Chadwick, and he related some interesting stories about the man. Chadwick was apparently a little nutty and a complete fanatic about science, to the extent that when he was held in a German prison camp during World War II, he managed to cajole his captors into allowing him to scrounge up parts from broken radios
so that he could attempt to do physics experiments.

Chadwick’s experiment worked like this. Suppose you perform two Rutherford-type alpha scattering measurements, first one with a gold foil as a target as in Rutherford’s original experiment, and then one with a copper foil. It is possible to get large angles of deflection in both cases, but as shown in figure 1, the alpha particle must be heading almost straight for the copper nucleus to get the same angle of deflection that would have occurred with an alpha that was much farther off the mark; the gold nucleus’ charge is so much greater than the copper’s that it exerts a strong force on the alpha particle even from far off. The situation is very much like that of a blindfolded person playing darts. Just as it is impossible to aim an alpha particle at an individual nucleus in the target, the blindfolded person cannot really aim the darts. Achieving a very close encounter with the copper atom would be akin to hitting an inner circle on the dartboard. It’s much more likely that one would have the luck to hit the outer circle, which covers a greater number of square inches. By analogy, if you measure the frequency with which alphas are scattered by copper at some particular angle, say between 19 and 20 degrees, and then perform the same measurement at the same angle with gold, you get a much higher percentage for gold than for copper.

In fact, the numerical ratio of the two nuclei’s charges can be derived from this same experimentally determined ratio. Using the standard notation $Z$ for the atomic number (charge of the nucleus divided by $e$), the following equation can be proved (example 6):

$$\frac{Z_{\text{gold}}^2}{Z_{\text{copper}}^2} = \frac{\text{number of alphas scattered by gold at 19-20°}}{\text{number of alphas scattered by copper at 19-20°}}$$

By making such measurements for targets constructed from all the elements, one can infer the ratios of all the atomic numbers, and since the atomic numbers of the light elements were already known, atomic numbers could be assigned to the entire periodic table. Ac-
cording to Mosely, the atomic numbers of copper, silver and plati-
num were 29, 47, and 78, which corresponded well with their posi-
tions on the periodic table. Chadwick’s figures for the same elements
were 29.3, 46.3, and 77.4, with error bars of about 1.5 times the fun-
damental charge, so the two experiments were in good agreement.

The point here is absolutely not that you should be ready to plug
numbers into the above equation for a homework or exam question!
My overall goal in this chapter is to explain how we know what we
know about atoms. An added bonus of describing Chadwick’s ex-
periment is that the approach is very similar to that used in modern
particle physics experiments, and the ideas used in the analysis are
closest related to the now-ubiquitous concept of a “cross-section.”
In the dartboard analogy, the cross-section would be the area of the
circular ring you have to hit. The reasoning behind the invention of
the term “cross-section” can be visualized as shown in figure 1. In
this language, Rutherford’s invention of the planetary model came
from his unexpected discovery that there was a nonzero cross-section
for alpha scattering from gold at large angles, and Chadwick con-
firmed Mosely’s determinations of the atomic numbers by measuring
cross-sections for alpha scattering.

Proof of the relationship between Z and scattering example 6
The equation above can be derived by the following not very rigor-
ous proof. To deflect the alpha particle by a certain angle requires
that it acquire a certain momentum component in the direction
perpendicular to its original momentum. Although the nucleus’s
force on the alpha particle is not constant, we can pretend that
it is approximately constant during the time when the alpha is
within a distance equal to, say, 150% of its distance of closest
approach, and that the force is zero before and after that part of
the motion. (If we chose 120% or 200%, it shouldn’t make any
difference in the final result, because the final result is a ratio,
and the effects on the numerator and denominator should cancel
each other.) In the approximation of constant force, the change
in the alpha’s perpendicular momentum component is then equal
to $F\Delta t$. The Coulomb force law says the force is proportional to
$Z/r^2$. Although $r$ does change somewhat during the time interval
of interest, it’s good enough to treat it as a constant number, since
we’re only computing the ratio between the two experiments’ re-
sults. Since we are approximating the force as acting over the
time during which the distance is not too much greater than the
distance of closest approach, the time interval $\Delta t$ must be propor-
tional to $r$, and the sideways momentum imparted to the alpha,$F\Delta t$, is proportional to $(Z/r^2)r$, or $Z/r$. If we’re comparing alphas
scattered at the same angle from gold and from copper, then $\Delta p$
is the same in both cases, and the proportionality $\Delta p \propto Z/r$ tells
us that the ones scattered from copper at that angle had to be headed in along a line closer to the central axis by a factor equaling \( \frac{Z_{gold}}{Z_{copper}} \). If you imagine a "dartboard ring" that the alphas have to hit, then the ring for the gold experiment has the same proportions as the one for copper, but it is enlarged by a factor equal to \( \frac{Z_{gold}}{Z_{copper}} \). That is, not only is the radius of the ring greater by that factor, but unlike the rings on a normal dartboard, the thickness of the outer ring is also greater in proportion to its radius. When you take a geometric shape and scale it up in size like a photographic enlargement, its area is increased in proportion to the square of the enlargement factor, so the area of the dartboard ring in the gold experiment is greater by a factor equal to \( \left( \frac{Z_{gold}}{Z_{copper}} \right)^2 \). Since the alphas are aimed entirely randomly, the chances of an alpha hitting the ring are in proportion to the area of the ring, which proves the equation given above.

As an example of the modern use of scattering experiments and cross-section measurements, you may have heard of the recent experimental evidence for the existence of a particle called the top quark. Of the twelve subatomic particles currently believed to be the smallest constituents of matter, six form a family called the quarks, distinguished from the other six by the intense attractive forces that make the quarks stick to each other. (The other six consist of the electron plus five other, more exotic particles.) The only two types of quarks found in naturally occurring matter are the “up quark” and “down quark,” which are what protons and neutrons are made of, but four other types were theoretically predicted to exist, for a total of six. (The whimsical term “quark” comes from a line by James Joyce reading “Three quarks for master Mark.”) Until recently, only five types of quarks had been proven to exist via experiments, and the sixth, the top quark, was only theorized. There was no hope of ever detecting a top quark directly, since it is radioactive, and only exists for a zillionth of a second before evaporating. Instead, the researchers searching for it at the Fermi National Accelerator Laboratory near Chicago measured cross-sections for scattering of nuclei off of other nuclei. The experiment was much like those of Rutherford and Chadwick, except that the incoming nuclei had to be boosted to much higher speeds in a particle accelerator. The resulting encounter with a target nucleus was so violent that both nuclei were completely demolished, but, as Einstein proved, energy can be converted into matter, and the energy of the collision creates a spray of exotic, radioactive particles, like the deadly shower of wood fragments produced by a cannon ball in an old naval battle. Among those particles were some top quarks. The cross-sections being measured were the cross-sections for the production of certain combinations of these secondary particles. However different the details, the principle was the same as that employed at the turn of the century: you smash things together and look at the fragments.
that fly off to see what was inside them. The approach has been compared to shooting a clock with a rifle and then studying the pieces that fly off to figure out how the clock worked.

**Discussion Questions**

**A** The diagram, showing alpha particles being deflected by a gold nucleus, was drawn with the assumption that alpha particles came in on lines at many different distances from the nucleus. Why wouldn’t they all come in along the same line, since they all came out through the same tube?

**B** Why does it make sense that, as shown in the figure, the trajectories that result in 19° and 20° scattering cross each other?

**C** Rutherford knew the velocity of the alpha particles emitted by radium, and guessed that the positively charged part of a gold atom had a charge of about +100e (we now know it is +79e). Considering the fact that some alpha particles were deflected by 180°, how could he then use conservation of energy to derive an upper limit on the size of a gold nucleus? (For simplicity, assume the size of the alpha particle is negligible compared to that of the gold nucleus, and ignore the fact that the gold nucleus recoils a little from the collision, picking up a little kinetic energy.)

**8.2.4 The structure of nuclei**

*The proton*

The fact that the nuclear charges were all integer multiples of $e$ suggested to many physicists that rather than being a pointlike object, the nucleus might contain smaller particles having individual charges of $+e$. Evidence in favor of this idea was not long in arriving. Rutherford reasoned that if he bombarded the atoms of a very light element with alpha particles, the small charge of the target nuclei would give a very weak repulsion. Perhaps those few alpha particles that happened to arrive on head-on collision courses would get so close that they would physically crash into some of the target nuclei. An alpha particle is itself a nucleus, so this would be a collision between two nuclei, and a violent one due to the high speeds involved. Rutherford hit pay dirt in an experiment with alpha particles striking a target containing nitrogen atoms. Charged particles were detected flying out of the target like parts flying off of cars in a high-speed crash. Measurements of the deflection of these particles in electric and magnetic fields showed that they had the same charge-to-mass ratio as singly-ionized hydrogen atoms. Rutherford concluded that these were the conjectured singly-charged particles that held the charge of the nucleus, and they were later named protons. The hydrogen nucleus consists of a single proton, and in general, an element’s atomic number gives the number of protons contained in each of its nuclei. The mass of the proton is about 1800 times greater than the mass of the electron.
The neutron

It would have been nice and simple if all the nuclei could have been built only from protons, but that couldn’t be the case. If you spend a little time looking at a periodic table, you will soon notice that although some of the atomic masses are very nearly integer multiples of hydrogen’s mass, many others are not. Even where the masses are close whole numbers, the masses of an element other than hydrogen is always greater than its atomic number, not equal to it. Helium, for instance, has two protons, but its mass is four times greater than that of hydrogen.

Chadwick cleared up the confusion by proving the existence of a new subatomic particle. Unlike the electron and proton, which are electrically charged, this particle is electrically neutral, and he named it the neutron. Chadwick’s experiment has been described in detail on p. 138, but briefly the method was to expose a sample of the light element beryllium to a stream of alpha particles from a lump of radium. Beryllium has only four protons, so an alpha that happens to be aimed directly at a beryllium nucleus can actually hit it rather than being stopped short of a collision by electrical repulsion. Neutrons were observed as a new form of radiation emerging from the collisions, and Chadwick correctly inferred that they were previously unsuspected components of the nucleus that had been knocked out. As described earlier, Chadwick also determined the mass of the neutron; it is very nearly the same as that of the proton.

To summarize, atoms are made of three types of particles:

<table>
<thead>
<tr>
<th>particle</th>
<th>charge</th>
<th>mass in units of the proton’s mass</th>
<th>location in atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>proton</td>
<td>+e</td>
<td>1</td>
<td>in nucleus</td>
</tr>
<tr>
<td>neutron</td>
<td>0</td>
<td>1.001</td>
<td>in nucleus</td>
</tr>
<tr>
<td>electron</td>
<td>−e</td>
<td>1/1836</td>
<td>orbiting nucleus</td>
</tr>
</tbody>
</table>

The existence of neutrons explained the mysterious masses of the elements. Helium, for instance, has a mass very close to four times greater than that of hydrogen. This is because it contains two neutrons in addition to its two protons. The mass of an atom is essentially determined by the total number of neutrons and protons. The total number of neutrons plus protons is therefore referred to as the atom’s *mass number.*

Isotopes

We now have a clear interpretation of the fact that helium is close to four times more massive than hydrogen, and similarly for all the atomic masses that are close to an integer multiple of the mass of hydrogen. But what about copper, for instance, which had an atomic mass 63.5 times that of hydrogen? It didn’t seem reasonable to think that it possessed an extra half of a neutron! The
solution was found by measuring the mass-to-charge ratios of singly-ionized atoms (atoms with one electron removed). The technique is essentially that same as the one used by Thomson for cathode rays, except that whole atoms do not spontaneously leap out of the surface of an object as electrons sometimes do. Figure n shows an example of how the ions can be created and injected between the charged plates for acceleration.

Injecting a stream of copper ions into the device, we find a surprise — the beam splits into two parts! Chemists had elevated to dogma the assumption that all the atoms of a given element were identical, but we find that 69% of copper atoms have one mass, and 31% have another. Not only that, but both masses are very nearly integer multiples of the mass of hydrogen (63 and 65, respectively). Copper gets its chemical identity from the number of protons in its nucleus, 29, since chemical reactions work by electric forces. But apparently some copper atoms have $63 - 29 = 34$ neutrons while others have $65 - 29 = 36$. The atomic mass of copper, 63.5, reflects the proportions of the mixture of the mass-63 and mass-65 varieties. The different mass varieties of a given element are called *isotopes* of that element.
Isotopes can be named by giving the mass number as a subscript to the left of the chemical symbol, e.g., \( ^{65}\text{Cu} \). Examples:

<table>
<thead>
<tr>
<th>( ^1\text{H} )</th>
<th>( ^4\text{He} )</th>
<th>( ^{12}\text{C} )</th>
<th>( ^{14}\text{C} )</th>
<th>( ^{262}\text{Ha} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>2</td>
<td>6</td>
<td>6</td>
<td>105</td>
</tr>
<tr>
<td>0</td>
<td>2</td>
<td>6</td>
<td>8</td>
<td>157</td>
</tr>
<tr>
<td>1+0 = 1</td>
<td>2+2 = 4</td>
<td>6+6 = 12</td>
<td>6+8 = 14</td>
<td>105+157 = 262</td>
</tr>
</tbody>
</table>

**self-check D**

Why are the positive and negative charges of the accelerating plates reversed in the isotope-separating apparatus compared to the Thomson apparatus?  

Chemical reactions are all about the exchange and sharing of electrons: the nuclei have to sit out this dance because the forces of electrical repulsion prevent them from ever getting close enough to make contact with each other. Although the protons do have a vitally important effect on chemical processes because of their electrical forces, the neutrons can have no effect on the atom’s chemical reactions. It is not possible, for instance, to separate \(^{63}\text{Cu}\) from \(^{65}\text{Cu}\) by chemical reactions. This is why chemists had never realized that different isotopes existed. (To be perfectly accurate, different isotopes do behave slightly differently because the more massive atoms move more sluggishly and therefore react with a tiny bit less intensity. This tiny difference is used, for instance, to separate out the isotopes of uranium needed to build a nuclear bomb. The smallness of this effect makes the separation process a slow and difficult one, which is what we have to thank for the fact that nuclear weapons have not been built by every terrorist cabal on the planet.)

**Sizes and shapes of nuclei**

Matter is nearly all nuclei if you count by weight, but in terms of volume nuclei don’t amount to much. The radius of an individual neutron or proton is very close to 1 fm (1 fm\(=10^{-15} \) m), so even a big lead nucleus with a mass number of 208 still has a diameter of only about 13 fm, which is ten thousand times smaller than the diameter of a typical atom. Contrary to the usual imagery of the nucleus as a small sphere, it turns out that many nuclei are somewhat elongated, like an American football, and a few have exotic asymmetric shapes like pears or kiwi fruits.

**Discussion Questions**

**A** Suppose the entire universe was in a (very large) cereal box, and the nutritional labeling was supposed to tell a godlike consumer what percentage of the contents was nuclei. Roughly what would the percentage be like if the labeling was according to mass? What if it was by volume?
The strong nuclear force cuts off very sharply at a range of about 1 fm.

A nuclear power plant at Cattenom, France. Unlike the coal and oil plants that supply most of the U.S.’s electrical power, a nuclear power plant like this one releases no pollution or greenhouse gases into the Earth’s atmosphere, and therefore doesn’t contribute to global warming. The white stuff puffing out of this plant is non-radioactive water vapor. Although nuclear power plants generate long-lived nuclear waste, this waste arguably poses much less of a threat to the biosphere than greenhouse gases would.

8.2.5 The strong nuclear force, alpha decay and fission

Once physicists realized that nuclei consisted of positively charged protons and uncharged neutrons, they had a problem on their hands. The electrical forces among the protons are all repulsive, so the nucleus should simply fly apart! The reason all the nuclei in your body are not spontaneously exploding at this moment is that there is another force acting. This force, called the strong nuclear force, is always attractive, and acts between neutrons and neutrons, neutrons and protons, and protons and protons with roughly equal strength. The strong nuclear force does not have any effect on electrons, which is why it does not influence chemical reactions.

Unlike electric forces, whose strengths are given by the simple Coulomb force law, there is no simple formula for how the strong nuclear force depends on distance. Roughly speaking, it is effective over ranges of \( \sim 1 \text{ fm} \), but falls off extremely quickly at larger distances (much faster than \( 1/r^2 \)). Since the radius of a neutron or proton is about 1 fm, that means that when a bunch of neutrons and protons are packed together to form a nucleus, the strong nuclear force is effective only between neighbors.

Figure q illustrates how the strong nuclear force acts to keep ordinary nuclei together, but is not able to keep very heavy nuclei from breaking apart. In q/1, a proton in the middle of a carbon nucleus feels an attractive strong nuclear force (arrows) from each of its nearest neighbors. The forces are all in different directions, and tend to cancel out. The same is true for the repulsive electrical forces (not shown). In figure q/2, a proton at the edge of the nucleus has neighbors only on one side, and therefore all the strong nuclear
1. The forces cancel. 2. The forces don’t cancel. 3. In a heavy nucleus, the large number of electrical repulsions can add up to a force that is comparable to the strong nuclear attraction. 4. Alpha emission. 5. Fission.

forces acting on it are tending to pull it back in. Although all the electrical forces from the other five protons (dark arrows) are all pushing it out of the nucleus, they are not sufficient to overcome the strong nuclear forces.

In a very heavy nucleus, q/3, a proton that finds itself near the edge has only a few neighbors close enough to attract it significantly via the strong nuclear force, but every other proton in the nucleus exerts a repulsive electrical force on it. If the nucleus is large enough, the total electrical repulsion may be sufficient to overcome the attraction of the strong force, and the nucleus may spit out a proton. Proton emission is fairly rare, however; a more common type of radioactive decay\(^1\) in heavy nuclei is alpha decay, shown in q/4. The imbalance of the forces is similar, but the chunk that is ejected is an alpha particle (two protons and two neutrons) rather than a single proton.

It is also possible for the nucleus to split into two pieces of roughly equal size, q/5, a process known as fission. Note that in addition to the two large fragments, there is a spray of individual

\(^1\)Alpha decay is more common because an alpha particle happens to be a very stable arrangement of protons and neutrons.
neutrons. In a nuclear fission bomb or a nuclear fission reactor, some of these neutrons fly off and hit other nuclei, causing them to undergo fission as well. The result is a chain reaction.

When a nucleus is able to undergo one of these processes, it is said to be radioactive, and to undergo radioactive decay. Some of the naturally occurring nuclei on earth are radioactive. The term “radioactive” comes from Becquerel’s image of rays radiating out from something, not from radio waves, which are a whole different phenomenon. The term “decay” can also be a little misleading, since it implies that the nucleus turns to dust or simply disappears – actually it is splitting into two new nuclei with an the same total number of neutrons and protons, so the term “radioactive transformation” would have been more appropriate. Although the original atom's electrons are mere spectators in the process of weak radioactive decay, we often speak loosely of “radioactive atoms” rather than “radioactive nuclei.”

Randomness in physics

How does an atom decide when to decay? We might imagine that it is like a termite-infested house that gets weaker and weaker, until finally it reaches the day on which it is destined to fall apart. Experiments, however, have not succeeded in detecting such “tick-ing clock” hidden below the surface; the evidence is that all atoms of a given isotope are absolutely identical. Why, then, would one uranium atom decay today while another lives for another million years? The answer appears to be that it is entirely random. We can make general statements about the average time required for a certain isotope to decay, or how long it will take for half the atoms in a sample to decay (its half-life), but we can never predict the behavior of a particular atom.

This is the first example we have encountered of an inescapable randomness in the laws of physics. If this kind of randomness makes you uneasy, you’re in good company. Einstein’s famous quote is “...I am convinced that He [God] does not play dice.” Einstein’s distaste for randomness, and his association of determinism with divinity, goes back to the Enlightenment conception of the universe as a gigantic piece of clockwork that only had to be set in motion initially by the Builder. Physics had to be entirely rebuilt in the 20th century to incorporate the fundamental randomness of physics, and this modern revolution is the topic of chapter 13. In particular, we will delay the mathematical development of the half-life concept until then.
8.2.6 The weak nuclear force; beta decay

All the nuclear processes we’ve discussed so far have involved rearrangements of neutrons and protons, with no change in the total number of neutrons or the total number of protons. Now consider the proportions of neutrons and protons in your body and in the planet earth: neutrons and protons are roughly equally numerous in your body’s carbon and oxygen nuclei, and also in the nickel and iron that make up most of the earth. The proportions are about 50-50. But, as discussed in more detail on p. 505, the only chemical elements produced in any significant quantities by the big bang\(^2\) were hydrogen (about 90%) and helium (about 10%). If the early universe was almost nothing but hydrogen atoms, whose nuclei are protons, where did all those neutrons come from?

The answer is that there is another nuclear force, the weak nuclear force, that is capable of transforming neutrons into protons and vice-versa. Two possible reactions are

\[
\text{n} \rightarrow \text{p} + \text{e}^- + \bar{\nu} \quad \text{[electron decay]}
\]

and

\[
\text{p} \rightarrow \text{n} + \text{e}^+ + \nu \quad \text{[positron decay]}
\]

(There is also a third type called electron capture, in which a proton grabs one of the atom’s electrons and they produce a neutron and a neutrino.)

Whereas alpha decay and fission are just a redivision of the previously existing particles, these reactions involve the destruction of one particle and the creation of three new particles that did not exist before.

There are three new particles here that you have never previously encountered. The symbol \(\text{e}^+\) stands for an antielectron, which is a particle just like the electron in every way, except that its electric charge is positive rather than negative. Antielectrons are also known as positrons. Nobody knows why electrons are so common in the universe and antielectrons are scarce. When an antielectron encounters an electron, they annihilate each other, producing gamma rays, and this is the fate of all the antielectrons that are produced by natural radioactivity on earth. Antielectrons are an example of antimatter. A complete atom of antimatter would consist of antiprotons, antielectrons, and antineutrons. Although individual particles of antimatter occur commonly in nature due to natural radioactivity and cosmic rays, only a few complete atoms of antihydrogen have ever been produced artificially.

The notation \(\nu\) stands for a particle called a neutrino, and \(\bar{\nu}\) means an antineutrino. Neutrinos and antineutrinos have no electric charge (hence the name).

\(^2\)The evidence for the big bang theory of the origin of the universe was discussed on p. 356.
We can now list all four of the known fundamental forces of physics:

- gravity
- electromagnetism
- strong nuclear force
- weak nuclear force

The other forces we have learned about, such as friction and the normal force, all arise from electromagnetic interactions between atoms, and therefore are not considered to be fundamental forces of physics.

**Example 7**

As an example, consider the radioactive isotope of lead $^{212}\text{Pb}$. It contains 82 protons and 130 neutrons. It decays by the process $n \rightarrow p + e^- + \bar{\nu}$. The newly created proton is held inside the nucleus by the strong nuclear force, so the new nucleus contains 83 protons and 129 neutrons. Having 83 protons makes it the element bismuth, so it will be an atom of $^{212}\text{Bi}$.

In a reaction like this one, the electron flies off at high speed (typically close to the speed of light), and the escaping electrons are the things that make large amounts of this type of radioactivity dangerous. The outgoing electron was the first thing that tipped off scientists in the early 1900s to the existence of this type of radioactivity. Since they didn’t know that the outgoing particles were electrons, they called them beta particles, and this type of radioactive decay was therefore known as beta decay. A clearer but less common terminology is to call the two processes electron decay and positron decay.

The neutrino or antineutrino emitted in such a reaction pretty much ignores all matter, because its lack of charge makes it immune to electrical forces, and it also remains aloof from strong nuclear interactions. Even if it happens to fly off going straight down, it is almost certain to make it through the entire earth without interacting with any atoms in any way. It ends up flying through outer space forever. The neutrino’s behavior makes it exceedingly difficult to detect, and when beta decay was first discovered nobody realized that neutrinos even existed. We now know that the neutrino carries off some of the energy produced in the reaction, but at the time it seemed that the total energy afterwards (not counting the unsuspected neutrino’s energy) was greater than the total energy before the reaction, violating conservation of energy. Physicists were getting ready to throw conservation of energy out the window as a basic law of physics when indirect evidence led them to the conclusion that neutrinos existed.
Discussion Questions

A In the reactions $n \rightarrow p + e^- + \bar{\nu}$ and $p \rightarrow n + e^+ + \nu$, verify that charge is conserved. In beta decay, when one of these reactions happens to a neutron or proton within a nucleus, one or more gamma rays may also be emitted. Does this affect conservation of charge? Would it be possible for some extra electrons to be released without violating charge conservation?

B When an antielectron and an electron annihilate each other, they produce two gamma rays. Is charge conserved in this reaction?

8.2.7 Fusion

As we have seen, heavy nuclei tend to fly apart because each proton is being repelled by every other proton in the nucleus, but is only attracted by its nearest neighbors. The nucleus splits up into two parts, and as soon as those two parts are more than about 1 fm apart, the strong nuclear force no longer causes the two fragments to attract each other. The electrical repulsion then accelerates them, causing them to gain a large amount of kinetic energy. This release of kinetic energy is what powers nuclear reactors and fission bombs.

It might seem, then, that the lightest nuclei would be the most stable, but that is not the case. Let’s compare an extremely light nucleus like $^4$He with a somewhat heavier one, $^{16}$O. A neutron or proton in $^4$He can be attracted by the three others, but in $^{16}$O, it might have five or six neighbors attracting it. The $^{16}$O nucleus is therefore more stable.

It turns out that the most stable nuclei of all are those around nickel and iron, having about 30 protons and 30 neutrons. Just as a nucleus that is too heavy to be stable can release energy by splitting apart into pieces that are closer to the most stable size, light nuclei can release energy if you stick them together to make bigger nuclei that are closer to the most stable size. Fusing one nucleus with another is called nuclear fusion. Nuclear fusion is what powers our sun and other stars.
Our sun’s source of energy is nuclear fusion, so nuclear fusion is also the source of power for all life on earth, including, this rain forest in Fatu-Hiva. The first release of energy by nuclear fusion through human technology was the 1952 Ivy Mike test at the Enewetak Atoll. This array of gamma-ray detectors is called GAMMASPHERE. During operation, the array is closed up, and a beam of ions produced by a particle accelerator strikes a target at its center, producing nuclear fusion reactions. The gamma rays can be studied for information about the structure of the fused nuclei, which are typically varieties not found in nature. Nuclear fusion promises to be a clean, inexhaustible source of energy. However, the goal of commercially viable nuclear fusion power has remained elusive, due to the engineering difficulties involved in magnetically containing a plasma (ionized gas) at a sufficiently high temperature and density. This photo shows the experimental JET reactor, with the device opened up on the left, and in action on the right.

8.2.8 Nuclear energy and binding energies

In the same way that chemical reactions can be classified as exothermic (releasing energy) or endothermic (requiring energy to react), so nuclear reactions may either release or use up energy. The energies involved in nuclear reactions are greater by a huge factor. Thousands of tons of coal would have to be burned to produce as
much energy as would be produced in a nuclear power plant by one kg of fuel.

Although nuclear reactions that use up energy (endothermic reactions) can be initiated in accelerators, where one nucleus is rammed into another at high speed, they do not occur in nature, not even in the sun. The amount of kinetic energy required is simply not available.

To find the amount of energy consumed or released in a nuclear reaction, you need to know how much nuclear interaction energy, $U_{\text{nuc}}$, was stored or released. Experimentalists have determined the amount of nuclear energy stored in the nucleus of every stable element, as well as many unstable elements. This is the amount of mechanical work that would be required to pull the nucleus apart into its individual neutrons and protons, and is known as the nuclear binding energy.

\[1H + ^2H \rightarrow ^3\text{He} + \gamma\]

The excess energy is almost all carried off by the gamma ray (not by the kinetic energy of the helium-3 atom). The binding energies in units of pJ (picojoules) are:

\[
\begin{align*}
^1\text{H} & \quad 0 \text{ J} \\
^2\text{H} & \quad 0.35593 \text{ pJ} \\
^3\text{He} & \quad 1.23489 \text{ pJ}
\end{align*}
\]

The total initial nuclear energy is $0 \text{ pJ} + 0.35593 \text{ pJ}$, and the final nuclear energy is $1.23489 \text{ pJ}$, so by conservation of energy, the gamma ray must carry off $0.87896 \text{ pJ}$ of energy. The gamma ray is then absorbed by the sun and converted to heat.

**self-check E**

Why is the binding energy of $^1\text{H}$ exactly equal to zero?  

Answer, p. 929

Figure 8 is a compact way of showing the vast variety of the nuclei. Each box represents a particular number of neutrons and protons. The black boxes are nuclei that are stable, i.e., that would require an input of energy in order to change into another. The gray boxes show all of the unstable nuclei that have been studied experimentally. Some of these last for billions of years on the average before decaying and are found in nature, but most have much shorter average lifetimes, and can only be created and studied in the laboratory.

The curve along which the stable nuclei lie is called the line of stability. Nuclei along this line have the most stable proportion
of neutrons to protons. For light nuclei the most stable mixture is about 50-50, but we can see that stable heavy nuclei have two or three times more neutrons than protons. This is because the electrical repulsions of all the protons in a heavy nucleus add up to a powerful force that would tend to tear it apart. The presence of a large number of neutrons increases the distances among the protons, and also increases the number of attractions due to the strong nuclear force.

8.2.9 Biological effects of ionizing radiation

Units used to measure exposure

As a science educator, I find it frustrating that nowhere in the massive amount of journalism devoted to nuclear safety does one ever find any numerical statements about the amount of radiation to which people have been exposed. Anyone capable of understanding sports statistics or weather reports ought to be able to understand such measurements, as long as something like the following explanatory text was inserted somewhere in the article:

Radiation exposure is measured in units of Sieverts (Sv). The