X. Modern Physics

Introduction

The frontiers of science lie along the fringes of what is already known and understood. New science is needed whenever current theories are unable to satisfactorily explain experimental data. More often than not, as modern physics developed in the first five decades of the twentieth century, those advances were the result of literal conflicting contentions. Unlike many of the exercises in previous units of this workbook, the conflicts were not so subtle: is light a particle or a wave? Is energy continuous or discrete? Can the paths of electrons in atoms be described exactly, as Newton's mechanics would predict?

The final answers to most of these questions lay in the irrefutable data of key experiments, such as Rutherford's gold foil experiment or the photoelectric effect; and the conclusions reached by theorists based on experimental data were rarely simple or intellectually satisfying, often defying 'common sense'. Therefore, this last set of conceptual exercises on modern physics are presented in a different format compared to previous chapters.

Four topics of modern physics are examined, along with the often strange or confusing results. Your task is to explain how the data resolved the issues raised by two (or more) conflicting contentions or theories being debated at the time. As with the other conceptual exercises in this workbook, the best approach is to begin with one or more basic hypotheses about the behavior of the universe, apply reason and logic to determine the natural consequences of those assumptions if true, and then compare those predictions to the actual behavior of Nature.

The problems presented here are of great historical significance in the advancement of physics. The answers to these questions are very likely to be found in textbooks or encyclopedia articles. However, one of the goals of this workbook was to help you to develop the important reasoning skills necessary to allow you to 'think through' problems with uncertain answers *without* having to rely on someone else to tell you what the answer should be based on expertise or authority.

So act like a scientist: read what others have to say; think about it and try to understand it; and then develop an explanation in your own thoughts and words. As Mr. Nicholls would surely say, it's better to be wrong and know why you are wrong, then to be right without really understanding *why* you are right.

1. The Structure of Atoms: Rutherford's Gold Foil Experiment

The Situation

John Dalton had presented a theory in 1803 suggesting that all matter was composed of indivisible particles, which he called atoms. Different substances had different physical and chemical properties, according to Dalton, because they were either composed of a particular type of atom (the elements); or they were composed of different types of atoms bonded together chemically in set ratio, as identified by the substance's chemical formula; or, if two substances had the same chemical formula, they were bonded to each in different patterns that result in different structures.

Atomic theory proved to be a very useful description of matter; however, for more than a century it remained only a theory with no direct evidence that these 'particles of matter' actually existed. It wasn't until 1905 that Einstein was able to show that the random 'jiggling' of small but visible particles such as pollen floating on water (called Brownian motion) could be explained by collisions of atoms with those particles.

However, even as experimental data was demonstrating the existence of atoms, other experiments were suggesting that atoms were not 'fundamental' but had structures of their own. In 1899, J.J. Thompson used data from cathode ray tube experiments (the forerunner of the television picture tube or computer monitor) to show that the basic unit of electricity was a particle, which he called the 'electron'.

The link between matter and electricity was already well-established by 1900; standard household AA batteries are a prime example of electricity being chemically generated from matter. So, therefore, what was the relationship between electrons and atoms? If electrons originate from atoms, then where are they stored? In other words, what is the structure of the atom?

The Experiment

Thompson proposed that the atom was composed of negatively-charged electrons embedded in some type of positively-charged, fluid-like substance. This model was alternatively called the 'raisin bun' or 'plum pudding' model. Of course, the accepted model today is the nuclear atom established by Ernest Rutherford: a very massive, very small positively charged nucleus is surrounded by electrons, which occupy more than 99% of the mostly empty volume of the atom. But, in fact, the nuclear model of the atom wasn't predicted: it was deduced.

Ernest Rutherford was an early researcher in the field of radiation. It was he who coined the terms 'alpha' and 'beta' radiation. Evidence collected by Rutherford showed that alpha particles seemed to be the very same thing as positively-charged helium atoms. (Today we would say that the alpha particle is the helium nucleus, which is two protons and two neutrons, stripped of its two electrons.) Beta particles, we now know, are just fast-moving electrons.

Rutherford probed the structure of the atom by firing high speed alpha particles at a very thin piece of gold foil. The gold atoms were known to be almost fifty times more massive than alpha particles. (Helium atoms have a relative mass of 4.0 units, while gold atoms have a relative mass of 197 units.) Cathode ray experiments also showed that electrons possessed very little of the atom's mass.

Rutherford found that most of the alpha particles passed through the gold foil with little deflection. However, some alpha particles were deflected at large angles, and a few seemed to recoil, returning back in the direction they came. This result was very surprising: he would later say it was like shooting an artillery shell at tissue paper and having it bounce off.

Conflicting Contentions

a) How did the gold foil experiment conclusively demonstrate that the raisin bun model of the atom was incorrect? How did it support the nuclear model of the atom?

b) Create at least one other possible atomic structure besides the nuclear and raisin-bun models and predict what would happen to the alpha particles in the gold foil experiment if your model was correct.

   Some possibilities: just because electrons are very light when they leave the atoms doesn't mean they can't be more massive inside the atom. Or, even if the positively charged particles (now called protons) possess most of the mass (along with the neutrons), suppose they didn't exist as a nucleus. Can you imagine other arrangements of the protons, neutrons and electrons in the atom?

Qualitative Reasoning

a) Suppose Rutherford had bombarded the gold foil with the negatively-charged, much lighter beta particles instead of alpha particles. What would he have observed experimentally?

b) Suppose Rutherford had used a much lighter atom instead of gold, such as aluminum (atomic mass = 27.0 units). What effect, if any, would that have on the experimental results?

c) Suppose alpha particles were massive but negatively-charged instead of positively charged. What effect, if any, would that have on the experimental results?

2. The Structure of Atoms: The Hydrogen Spectrum and Bohr's model

The Situation

Rutherford proposed an atomic structure consisting of a very small, massive, positively-charged nucleus orbited by very light, negatively-charged electrons in an atom that was mostly empty space. With the exception of the charges, Rutherford's model is very similar to our massive sun orbited by a few planets in a solar system that is mostly empty space.

But the charges which electrons and the nucleus possess (and which the sun and planets don't) are extremely significant in the formulation of the model. Even as he proposed it, Rutherford knew the model had problems. In the theory of classical electromagnetism, an accelerating charge must radiate energy. The next set of questions demonstrate why this principle meant that Rutherford's model had to be incorrect.

Can You Explain This?

a. In Rutherford's simplest version of his atomic model, electrons orbited the nucleus at a constant speed in circular orbits. Can you explain how the electrons must have an acceleration under these conditions?

b. According to electromagnetic theory, accelerating charges must radiate energy. Therefore, orbiting electrons must crash into the nucleus, destroying the atom. Can you explain why?

c. Atoms do radiate energy, often in the form of visible light, which is called the atom's spectrum. We will consider atomic spectra shortly, but one key observation about atomic spectra is that the frequencies (or colors) of an atom's spectrum are characteristic of that atom. If electrons did crash into the nucleus and yet somehow resurrect themselves to begin orbiting again, we would observe a continually changing spectrum with the frequencies of light increasing from the red end toward the violet end of the visible spectrum as the electron spiraled in toward the nucleus. Why would that happen?

d. We can avoid the problem of energy loss through radiation if the electrons were at rest in the atom. What is the problem with that assumption?

The Experiment

A solution to Rutherford's planetary model of the atom was developed by one of his students, Niels Bohr, in 1911. A spectrum is a set of electromagnetic frequencies radiated by matter. Scientists had observed two types of spectra emitted by matter. A continuous spectrum, such as that produced by an incandescent lightbulb, shows all colors of light and occurs when matter absorbs thermal energy. Heated objects are able to glow with different colors (yellowish, bluish, etc.) as a result of the relative amount of radiation emitted at different frequencies of visible light.

On the other hand, if a rarified gas composed of a single type of atom is bombarded by high speed electrons, then the spectrum produced by the atom is discrete. Instead of showing all colors of light, only a very few, well-defined frequencies of light are produced by the glowing gas. Instead of literally tens of thousands of frequencies produced as with a continuous spectrum, there may be only a dozen or fewer colors of light in a discrete spectrum, which is also called a line spectrum. The colors of neon lights are often the result of the spectra produced by this method. For example, neon gas glows orange because most of its spectral lines are in the red, orange, or yellow region of the visible spectrum.

The line spectra of atoms helped Bohr develop his modified planetary model of the atom. Bohr began with Rutherford's assumption that electrons orbit the nucleus in circular orbits, but he added the condition that electrons were confined to a particular set of well-defined energy levels. According to his model, electrons could only absorb or release an amount of energy equal to the difference in energy between two particular levels. Starting with an equation he developed based on his ideas and applying the laws of classical mechanics to Coulomb's law of electric force, he was able to calculate the radius, speed, and energy of the electron for each of the allowed energy levels of the hydrogen atom.

The predictions of his model exactly matched the know frequencies of the hydrogen spectrum. He also demonstrated from his model that an experimental constant calculated from the hydrogen spectrum was actually just a combination of other well-established constants (the speed of light, the mass of the electron, etc.) that resulted naturally from the application of Newton's laws to the hydrogen atom when his ad hoc equation was included as a starting point. Therefore, even though Bohr's model only worked correctly for the hydrogen atom and other ions with only a single electron (such as He +), the model must contain some elements of a more accurate theory explaining the behavior of atomic and subatomic particles.

Can You Explain This?

e. How did Bohr's revised planetary model prevent the eventual destruction of the atom inherent in Rutherford's model?

Qualitative Reasoning

Recall that red light has a lower frequency and longer wavelength than blue light. According to a theory proposed by Max Planck, light of longer wavelength and lower frequency also possess less energy than higher frequency, shorter wavelength radiation. Therefore, red light has less energy than blue light.

Suppose an atom has only three energy levels. Also, suppose the lowest energy level is assigned a value of 0.0, the second energy level has a value of 10.0, and the highest energy level has a value of 15.0 relative energy units. Notice that the second and third energy levels have only half the spacing of the first and second energy levels.

a. How many transitions are allowed in this atom? That is, how many lines will be seen in this atom's spectrum?

b. Which transition will cause the line farthest toward the blue end of the spectrum? Which transition will result in the line closest to the red end of the spectrum?

c. How would the spectrum change compared to the original one if all of the energy levels were shifted by 5 energy units: that is, the first, second, and third energy levels are now 5.0, 15.0, and 20.0 energy units, respectively? How would the location of the lines in the spectrum or their relative positions change, if at all?

d. How would the spectrum change compared to the original one if all of the original energy levels were cut in half, so that the first, second, and third energy levels are now 0.0, 5.0, and 7.5 energy units, respectively? How would the location of the lines in the spectrum or their relative positions change, if at all?

e. How would the spectrum change compared to the original one if only the second energy level was increased by two units, so that the first, second, and third energy levels are now 0.0, 12.0, and 15.0 energy units, respectively? Once again, how would the location of the lines in the spectrum or their relative positions change, if at all?

3. Particles and Waves: The Photoelectric Effect

The Situation

Sir Isaac Newton provided some of the first definitive experimental evidence concerning the nature and behavior of light, and he believed that light was probably a particle. A contemporary of his, Christian Huygens, put forth a theory that light was a wave. The evidence for either model was inconclusive until 1801 when Thomas Young presented the first experimental evidence irrefutably supporting a wave model of light by showing that light could be diffracted, since that property is unique to waves. In 1862, James Clerk Maxwell presented a mathematical theory showing how light could be an electromagnetic wave. In his model, light was a natural form of radiation that resulted from electric charges vibrating with a regular frequency, and the frequency of the electromagnetic wave was the same as the frequency of the oscillating charge.

Ironically, the experiments conducted by Heinrich Hertz confirming the predictions of Maxwell's model, and therefore the wave theory of light, also laid the foundations for the resurrection of the particle model of light. Hertz noticed when a voltage was applied to two pieces of metal that was *almost* but not quite large enough to cause a spark to jump between the two surfaces, that shining a light on the metal surface *without* changing the voltage could induce the spark to jump anyway. This interaction between light and electricity was dubbed the photoelectric effect. It is still the basis for the 'electric eye' that automatically opens and closes sliding doors.

The photoelectric effect was studied by a number of scientists, with unexpected results. An apparatus was created which allowed for an adjustable applied voltage and the careful measurement of current. Different types of metals such as zinc and aluminum could be tested in the apparatus. Also, different frequencies (or colors) of light with varying intensity could be illuminated onto their surfaces. Here are the confusing facts that were observed from the careful study of the photoelectric effect:

1. If the photoelectric effect occurred, then as the intensity of light increased, the flow of electricity also increased.

2. Each type of metal had a unique 'cut off' frequency: all colors of light with a lower frequency or longer wavelength would not demonstrate the photoelectric effect; all colors of light with a higher frequency or shorter wavelength would produce the photoelectric effect in the metal.

3. If the effect occurred, then even a very low light intensity would cause a current.

4. No matter what the intensity (and therefore, the total energy) of light below the cut-off frequency, the photoelectric effect would *not* occur. In other words, a weak blue pen light

shining on the metal might cause the photoelectric effect, but a high intensity heat lamp shining on the same metal would not.

Conflicting Contentions

a. Use logic and reason to predict the behavior of the photoelectric effect apparatus based on Maxwell's wave model of light. Could that model explain the observations made by scientists?

In response to the data and observations obtained from the photoelectric effect, Einstein resurrected a particle model of light. He proposed that light was produced and absorbed as localized bundles of energy, which he called photons. His photons retained a characteristic frequency, which until then was only associated with waves. He also retained Planck's idea that the energy of a photon was proportional to its frequency, so that a photon with twice the frequency of another also had twice the energy.

b. Use logic and reason to predict the behavior of the photoelectric effect apparatus based on Einstein's particle model of light.

Qualitative Reasoning

Einstein's explanation of the unique cut-off frequencies for different metals was that the force of attraction between atoms and the surface electrons depended on the type of atom. He called the minimum amount of energy needed to free the electron from the metal surface the *work function*.

a. Suppose aluminum had a stronger force of attraction to its surface electrons compared to zinc. Which metals would have the larger work function? Which metal would have the higher cut-off frequency?

Electrons are released from the negative electrode and travel through a vacuum tube to the positive electrode. Normally, only the negative electrode is illuminated. Assume both the positive and negative electrodes are flat surfaces made of the same metal. Also assume that current flow is the result of the photoelectric effect only — the voltage on the electrodes is not large enough to cause current to flow otherwise.

b. What would be the effect on the current if both plates are illuminated at the same time with the same intensity of light above the cut-off frequency?

c. What would be the effect on the current if only the positive electrode was illuminated and not the negative electrode?

Can You Explain This?

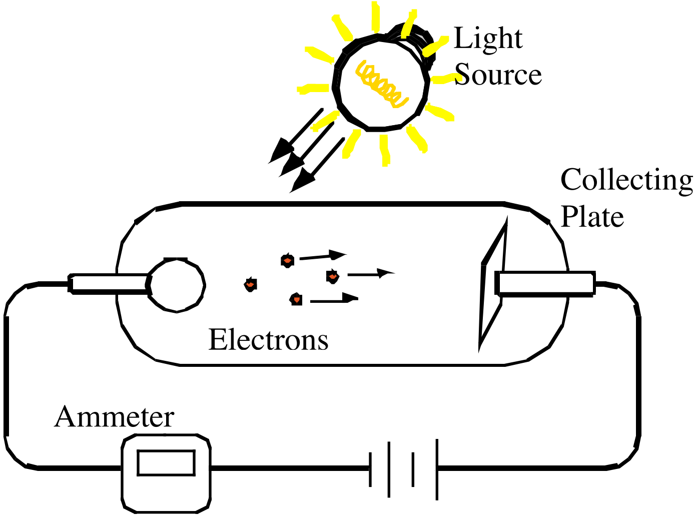
The photoelectric effect apparatus also allowed an experimenter to slow down and stop the flow of electricity by applying a 'stopping voltage'. The collecting plate could be made more and more negative until the Coulomb force repelled all electrons trying to reach the surface, so the current fell to zero.

a. Data shows that increasing the intensity of the illumination does not change the stopping voltage for a particular frequency of light. From this observation, scientists concluded that increasing the intensity of the light increased the *number* of electrons but *not* the average kinetic energy of the electrons produced by the effect. Can you explain why this is a valid conclusion?

b. Data also shows that illumination of the negative electrode with a higher frequency of light does require a larger stopping voltage. Can you explain why?

Hint: Einstein's work function is the minimum energy needed to free an electron from a metal's surface. The amount of energy of a photon is proportional to its frequency. What happens to the rest of the photon's energy after it is absorbed by the electron?

FYI: Einstein proposed the Special Theory of Relativity, explained the photoelectric effect, and demonstrated the real existence of atoms through Brownian motion in a single year: 1905. When he was awarded the Nobel Prize in physics, his development of Special Relativity was not mentioned in the award announcement — but his work on the photoelectric effect was.



4. Particles and Waves: Experiments and Orbitals

The Situation

The model of the photon as a particle of light was well-established when Bohr proposed his new model of the atom. From this perspective, an electron absorbs or emits a photon of light equal to the energy difference between two allowed orbits, but Bohr did not have an explanation *why* electrons could possess only certain energies in an atom. Furthermore, his model only worked for the hydrogen atom with its single electron, or other atoms that were stripped of all electrons but one.

In 1923, a French graduate student, Louis de Broglie, proposed a strange and radical theory in his doctoral thesis that almost prevented him from getting a Ph.D. He reasoned that if a light wave could also have a particle-like nature as a photon, then perhaps a particle such as an electron could also have a wave-like nature. He proposed that the wavelength associated with a particle was related to its momentum (p) which is the product of its mass and velocity, p = mv. According to de Broglie, a particle with a greater the momentum also has a shorter the wavelength (_). The conversion between momentum and wavelength was related to Planck's constant (h) , which, you may recall, was first proposed as the conversion factor between the energy and frequency of light.

Planck's Relation:	E = h f
de Broglie's Relation	mv = p = h/□

Qualitative Reasoning

a) If moving, massive particles have a wave nature, then it is appropriate to ask why we don't normally identify the wave nature of matter in our everyday lives. The answer lies in the size of Plank's constant, which in SI units has a magnitude around 10-34 units. Suppose a 1000 kg car is speeding along at 20 m/s. How long is its wavelength ? Use de Broglie's relation, rewritten as

   Wavelength = Plank's constant / momentum

   Compare the wavelength of the car estimated from de Broglie's relation (which will be in meters) to the diameter of the atom, around 10^{-10} meters, and the diameter of the nucleus, 10^{-15} meters. Which of these three quantities is the smallest?

b) An electron has a mass of around 10^{-31} kg. Given that Planck's constant is around 10^{-34} units and assuming an electron wavelength on the same order as the diameter of an atom, 10^{-10} meters, how fast would it have to be orbiting the nucleus? Make an estimate using the relation:

   Speed = Planck's constant / (mass x wavelength)

   Given that electrons exiting the electron gun at the back of a computer monitor are moving over 1000 m/s, is this speed of the electron in the atom unreasonable?

Can You Explain This? Electrons and crystals

In 1925, two years before de Broglie's hypothesis, Clinton Davisson and co-workers observed some puzzling results when high speed electrons passed through a nickel crystal. Instead of detecting the electrons in a typical 'scatter pattern' one might expect of particles (for example, the type of pattern observed if BBs were shot at a solid target), the electrons clustered at certain distinct and regularly spaced positions on the detection screen. Scientists soon realized that Davisson's experiments were direct evidence that electrons did possess a wave nature.

What wave-like property were the electrons displaying?

HINT: The nickel atoms were equally spaced in the crystal, and the electron wavelength predicted by de Broglie, based on the mass and speed of the electrons, was on the same order as the spacing. Also, it may help to review the four properties of light waves in the optics unit.

Qualitative Reasoning: A Basis for the Bohr Model

Once de Broglie's hypothesis was accepted by scientists of his day, 'wave-particle duality' became the basis for the 'quantum theory' developed by Bohr and others.

Bohr's equation that allowed him to predict the spectrum of the hydrogen atom could be reinterpreted as the equation that would result if electrons moved around the atom as circular standing waves. An electron would possess a certain wave amplitude at a certain point in space. As it moved around the atom, the amplitude would increase and decrease. The standing wave condition simply meant that once the wave returned to its original starting point, it must possess the same wave amplitude as when it started out.

a) Why is this a reasonable condition? What would be the implications if the amplitude was different at the starting point after one complete revolution of the electron?

Consider the first standing wave condition, when the circumference of the orbit is equal to one complete wavelength of the electron. Now imagine the electron being moved farther from the nucleus until it reaches the second standing wave condition, when the circumference of its orbit is equal to exactly two standing waves.

b) How does the energy of the electron in the second orbit compare to its energy in the first? Should it be more, less, or the same as the energy in the first orbit, and why?

c) Will the wavelength of the electron in the second orbit be longer, shorter, or the same as the wavelength of the electron in the first orbit? Be sure to give a reasonable explanation for your answer.

d) Will the distances between consecutive electron orbits be constant? Why or why not?

e) What experimental evidence suggests that the change in energy of the electron from orbit to orbit is not the same? Hint: examine the line spectra produced by the hydrogen atom.

Can You Explain This? Orbitals

One reason why de Broglie's hypothesis was both radical and controversial was that no one at the time had a good explanation for what the 'wave' actually represented. After the reality of a wave nature was firmly established, Max Born proposed that the amplitude of the electron wave represented a statistical probability.

Heisenberg's uncertainty principle implies that one cannot know both the location of an electron in an atom and its momentum (that is, where it is going) both at the same time. This principle combined with Born's hypothesis leads to the inevitable conclusion that electron orbits are not possible — or, at least are not detectable. Therefore, the concept of an electron orbit is replaced with that of an orbital. A visual representation of an orbital is a fuzzy cloud of a certain shape in a region of space. The more dense the cloud, the more likely one would find the electron there, if a person was to look for it. The density of the cloud is related to the square of de Broglie's electron wave amplitude.

> How does the concept of an orbital satisfy Rutherford's problem of continual radiation of energy resulting in an unstable atom?