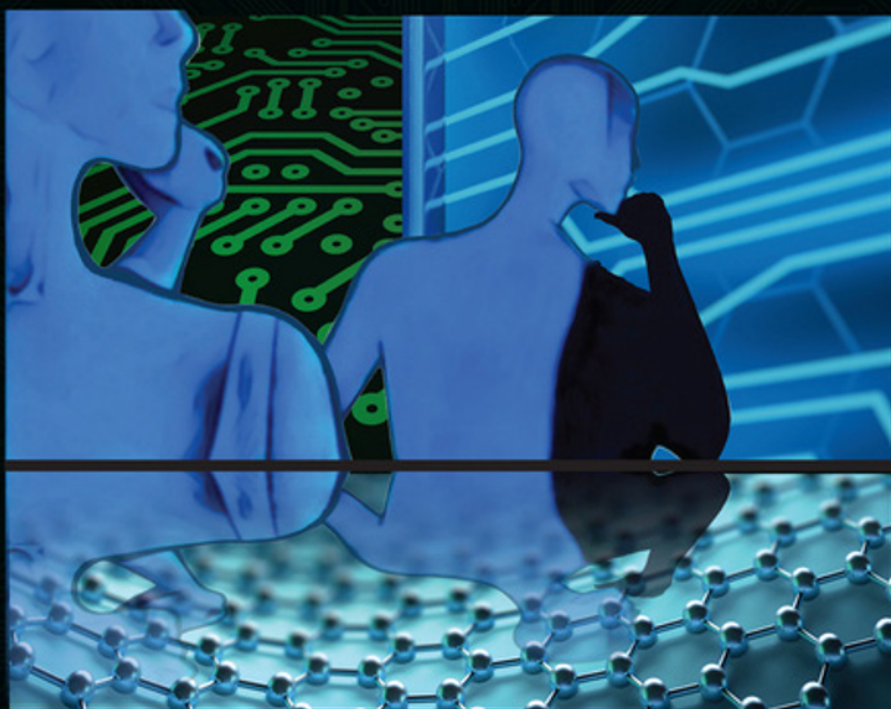


SEMICONDUCTOR BASICS

A qualitative, non-mathematical explanation of
how semiconductors work and how they are used



George Domingo

WILEY

Semiconductor Basics

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A Qualitative, Non-mathematical Explanation of How
Semiconductors Work and How They Are Used

George Domingo
Berkeley
CA, USA

WILEY

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To my family for their love and support

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I want to acknowledge the efforts and support of the Wiley editors who made the text more readable and clearer.

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Finally, I would like to mention Dr. Gerry Hanh, who upon reading the first chapters that I wrote as a hobby, insisted I send them to Wiley, resulting in the book you have now in your hands.

Introduction

A couple of years ago, I was asked to give a talk to the Rotary club in Terrassa, a city about 20 miles northwest of Barcelona. They asked me to divide the talk into three parts: 15 minutes biographical, 15 minutes on my technical work, and 15 minutes about NASA. The first and last 15 minutes of the talk went well, but the technical explanation about how infrared detectors work was disappointing. Yes, they did understand the uses and applications of infrared detectors, my technical work with the astronomical observatories of NASA, but the explanation on how infrared detectors work was not as clear and instructive as I had hoped. I was then and I am now convinced that any educated person can understand how semiconductor devices work. This talk two years ago was my motivation to start writing this book.

Semiconductors are the basis for almost all modern electronic devices. For many people semiconductors are a mysterious material that somehow has taken over modern electronics. In the same way that we understand the concept of god and creatures, but semi-gods are confusing, most of us have an idea of what a conductor (electricity flows through it) and an insulator (it doesn't) are, but what the heck is a semiconductor? Furthermore, the prevalent material used for fabricating semiconductors is pure silicon, the second most common element found on earth (28%) after oxygen (47%). Why not use aluminum, the next common element (8%), or strontium or some other exotic and more classy material? Why do we use semiconductors instead of conductors? Don't we want electrons to move freely through our devices?

I attempt here to explain why and how semiconductors work in a form that any educated person can understand. Every chapter explains the subject in a qualitative way, with drawings, photographs, and figures, and very simple relationships (I hate to use the word "equations" here). At the end of most chapters I add appendices where I include some mathematical formulation with the relevant equations for those who would enjoy looking a little deeper into the subject. I do not try to prove these equations; my purpose is not to formally present to you the physics of semiconductors (there are many excellent books that do that) but just to show what the results are. In this book you have to take these results on faith.

Unless you are very allergic to math, do not skip the appendices. There are interesting concepts that amplify the understanding of how semiconductors work. Don't worry; there are no problems at the end of each chapter and no suggested quizzes.

First, I explain what a semiconductor is, the different types we use, and how they are different from conductors and insulators. Next, I explain the key devices that can be constructed using the semiconductor materials: diodes, passive element, and transistors. I talk about the integrated circuits, how we build them and the larger electronic components, and finally what can we expect in the future (Chapter 15). I interrupt the “theoretical” flow with chapters devoted to applications (Chapters 4, 7, and 9) that can be understood with just the concepts I have covered in the previous chapters.

This book can be read in different ways. If you are interested in understanding how transistors work, then you should read, in succession, Chapters 1, 2, 3, 5, and 8. You cannot skip any of them or read them in a different order unless, of course, you are already familiar with some of the previous topics. Chapter 6 explains the basic electrical components (resistors, capacitors, and inductors) which we need to understand how we build useful semiconductor circuits. I try to discuss some of the semiconductor applications as soon as I cover the relevant theory. Just by understanding the concept of energy levels and energy bands (Chapters 2 and 3) you can grasp how infrared detectors work (Chapter 4). You do not even need to know what resistors or capacitors are. Similarly, after Chapter 6 you can understand the different circuit applications we can fabricate with diodes (Chapter 7) and after Chapter 8 I explain how we use transistors (Chapter 9). The next chapters, integrated circuit fabrication (Chapter 10) and logic circuits (Chapter 11), can be read separately at any time, although knowing how a transistor works will help. Yes, I talk about semiconductor devices, but you do not need to know the physics of how they work to understand these two chapters.

Next, in Chapter 12, I explain large semiconductor components like multiplexers and memories. In Chapter 13, optoelectronics, I cover lasers and LEDs, and I end the book by discussing some simple concepts related to computer architectures and liquid crystal displays in Chapter 14. I speculate about the future in Chapter 15.

The objective of this book is to explain to the layman how semiconductors work and how they are used. I hope I have succeeded. If you have any comments or if you find errors, inconsistencies, unclear, incomplete or confusing figures or explanations, please let me know by emailing me at semiconductorbasics@hotmail.com.

I had a professor who taught me relativistic quantum mechanics, one of the hardest courses I ever took, who used to say, “Everything mathematical is trivial.” I agree with him, not 100%, but partially so. The truth is that we scare the hell out of young students by making them believe that mathematics, physics, and engineering are difficult. Wrong! Math is not difficult and physics, and all its derivatives, is fascinating.

March 2019

*So please enjoy.
George Domingo*

1

The Bohr Atom

OBJECTIVES OF THIS CHAPTER

To understand how semiconductors work and how they are used, we need to be familiar with the concept of allowed energy levels first proposed by Niels Bohr. How Bohr came up with the idea of a planetary model of the atom is very interesting. Science is a continuum: one observation leads to a hypothesis that leads to a theory that leads to more observations, and so on. Bohr did not come up with his model of the atom out of the blue – no apple fell on his head. A lot of observations and theories going back to the 1700s were proposed before he put them together in a now well-known theory.

In this chapter, I discuss the experimental evidence and the scientific observations that led to Bohr's planetary model of the atom and the discrete energy levels it postulated. This brief journey will help us understand the significance of the Bohr atom that explains the strange behavior of light spectra.

1.1 Sinusoidal Waves

Before I start, I want to clarify the terms used to define a wave, which I use in the next few sections, and the relations between these terms (see Figure 1.1). There are four variables that we use to define any sinusoidal wave:

- The *amplitude*, A : How high each peak of the wave is in relation to the middle, its zero value.
- The *frequency*, f : The number of ups and downs in the wave in a given time. The units are Hertz or number of ups and downs per second: a number/s.
- The *wavelength*, λ : The distance between two peaks, in meters (m), centimeters (cm), or micrometers (μm).
- The *wave number*, ν (the Greek letter nu, not the letter ν): The reciprocal of the wavelength. Some properties of waves are better expressed by the reciprocal of the wavelength. The units are therefore $1/\text{m}$ or m^{-1} , or cm^{-1} , or μm^{-1} .

The last three variables are related by the velocity of the wave. *Velocity* is the distance that a moving object covers during a fixed amount of time, so the velocity v (this is now the letter v) has units of meters per second (m/s). The key relationships are

$$f = \frac{v}{\lambda} \quad \text{or} \quad \lambda = \frac{v}{f} \quad (1.1)$$

and the wave number – the reciprocal of the wavelength – is

$$v = \frac{1}{\lambda} = \frac{f}{v} \quad (1.2)$$

For example, suppose that Figure 1.1 represents a sound wave. The velocity of sound in air is 343 m s^{-1} . Take a look at Figure 1.1:

- The figure shows 5 cycles in 0.001 seconds, which means the frequency is 5000 cycles per second or $f = 5000 \text{ Hz}$, (where Hz, Hertz is the unit for frequency) which happens to be the middle of our hearing range.
- The wavelength is the velocity divided by the frequency, or $\lambda = 343 (\text{m s}^{-1})/5000 (1 \text{ s}^{-1}) = 0.069 \text{ m}$ or 6.9 cm. Notice that the seconds cancel out, and therefore the units are in meters or centimeters.
- The wave number is $v = 1/0.069 \text{ m} = 14.5 \text{ m}^{-1}$.

As much as possible, I use the metric system of units (MKS, meter, kilogram, second). I have always found it very annoying when books keep changing the unit system. When necessary, I will give you the equivalents.

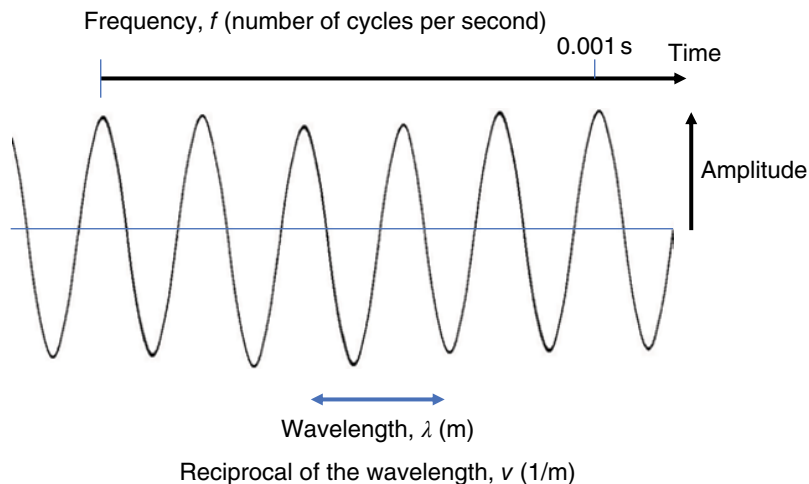


Figure 1.1 A sinusoidal wave is described in several ways: frequency, wavelength, and reciprocal of the wavelength plus its amplitude.

Now we are ready to dive into the pre-history of the Bohr atom and understand how Dr. Bohr came up with his famous model.

1.2 The Case of the Missing Lines

To explain how semiconductors work, we start with the Bohr atom. Most readers are familiar with Bohr's planetary model of the atom. How Bohr came up with this model is a very interesting scientific historical path involving many famous scientists of the nineteenth and early twentieth centuries. Science goes one step at a time.

William Wollaston (1766–1826), shown on the left in Figure 1.2, was an English chemist who discovered a couple of atomic elements, including palladium and rhodium. Very early in the 1800s, he built the first spectrometer. Wollaston focused the light from the sun through a prism and, to his surprise, found black lines partitioning the spectra (Figure 1.3). What the heck was going on?

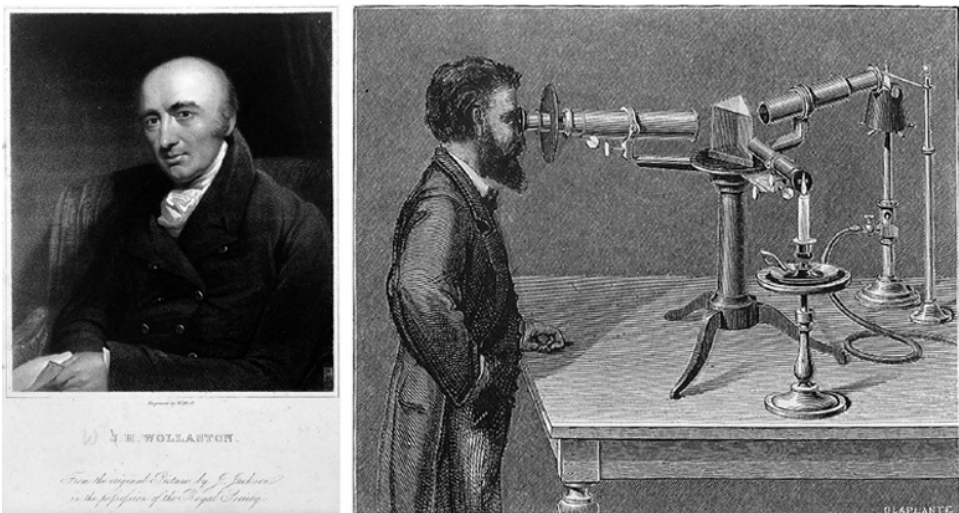


Figure 1.2 William Wollaston (left) looked at the sun's light through a prism and was the first to observe the missing lines. *Source:* <https://library.si.edu/image-gallery/73731>. Joseph von Fraunhofer (right) studied the missing lines with his spectrometer and named them A–K, where Hz, Hertz is the unit for frequency. <https://www.krueess.com/en/campus/spectroscopy/history-of-spectroscopy/>

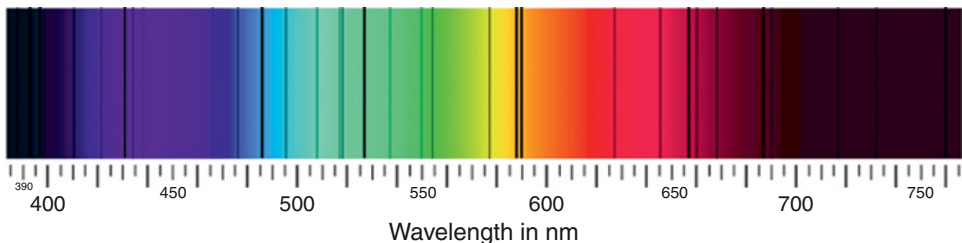


Figure 1.3 The sun's spectrum through a prism shows dark lines: wavelengths of light that seem to have disappeared. *Source:* <https://www.krueess.com/en/campus/spectroscopy/history-of-spectroscopy/>. (See insert for a color representation of this figure)

Suppose you are roasting a chicken and carefully watching the dial of a digital thermometer inserted in the chicken's breast as the temperature increases from room temperature, 24°C (degrees Celsius), to 80°C , the recommended internal temperature of a well-cooked chicken breast. As the temperature increases, suddenly the thermometer jumps from 39°C to 41°C , then from 56°C to 58°C , and finally from 66°C to 68°C . You wonder what is wrong with the thermometer: why don't the temperatures 40 , 57 , and 67°C show up on the dial? They don't seem to exist. You buy a new thermometer, just to be sure, and find that exactly the same temperature values are missing. A third thermometer gives the same results. You place the same thermometers in soup, and the thermometers are well behaved, showing in succession the values 39 , 40 , and 41°C . So, the thermometers work. The missing temperatures are no coincidence. There is something in that chicken that makes the temperature jump from one value to another without passing through the one in the middle.

Well, that was probably Wollaston's initial reaction. What separated the colors? Was the instrument lens dirty? He even considered the possibility that there were natural boundaries between certain colors. But why didn't these black boundaries appear when he pointed the spectrometer at a white light?

German physicist Joseph von Fraunhofer (1787–1826), on the right in Figure 1.2, studied these dark lines of the sun's spectrum in much more detail and actually named the missing lines with the letters A–K (not too imaginative; ancient astronomers would have found much more attention-grabbing names).

1.3 The Strange Behavior of Spectra from Gases and Metals

The next step was to take a look at the emission of light from several metals, gases, and stars. The observers soon found that each element has its own unique set of lines. Take a look at Figure 1.4. White light through a prism generates a continuous spectrum (top diagram); no colors are missing. If we heat a gas until it glows and pass the light through the same prism (middle diagram), only certain lines are projected onto the screen: the rest of the

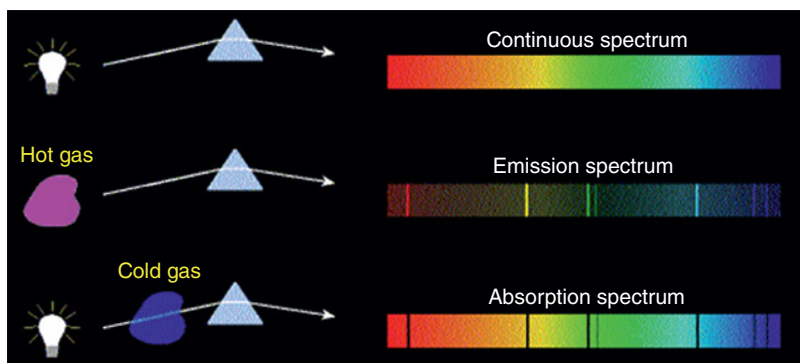



Figure 1.4 The spectrum from any gas shows similar but different missing lines (middle image), but when the same gas is hot and emits light, only the lines that were black before are now visible (lower image). Source: <https://quizlet.com/102018176/astronomy-4-spectroscopy-flash-cards/>. (See insert for a color representation of this figure)

spectrum has disappeared. But if we do it differently – that is, if we have the cold gas between the white light and the prism (bottom diagram) – then the full spectrum appears, except for *exactly* the same lines that were visible in the previous spectral measurement. The superposition of the middle and lower spectra is equal to the spectrum of the white light at the top. The gas, when cold, absorbs the same specific light waves as the ones it emits when it glows. What a coincidence! And this happens with *all* gases and materials. The lines are at different frequencies, but all of them have lines.

1.4 The Classifications of Basic Elements

In 1766, an English aristocrat named Henry Cavendish (1731–1810) was the first to recognize hydrogen as an element, that is, a unique substance and an integral part of the water molecule. By the time Dmitri Mendeleev (1834–1907) published his now ubiquitous periodic table of the elements in 1869, 100 years later, it was already known that hydrogen was the lightest of all the known elements (at that time, 60 elements were known). Mendeleev (Figure 1.5) reordered the known elements by their relative atomic weight, with hydrogen as 1.



I H 1.01	II	III	IV	V	VI	VII	VIII		
Li 6.94	Be 9.01	B 10.8	C 12.0	N 14.0	O 16.0	F 19.0			
Na 23.0	Mg 24.3	Al 27.0	Si 28.1	P 31.0	S 32.1	Cl 35.5			
K 39.1	Ca 40.1		Ti 47.9	V 50.9	Cr 52.0	Mn 54.9	Fe 55.9	Co 58.9	Ni 58.7
Cu 63.5	Zn 65.4			As 74.9	Se 79.0	Br 79.9			
Rb 85.5	Sr 87.6	Y 88.9	Zr 91.2	Nb 92.9	Mo 95.9		Ru 101	Rh 103	Pd 106
Ag 108	Cd 112	In 115	Sn 119	Sb 122	Te 128	I 127			
Ce 133	Ba 137	La 139		Ta 181	W 184		Os 194	Ir 192	Pt 195
Au 197	Hg 201	Tl 204	Pb 207	Bi 209					
			Th 232		U 238				

Figure 1.5 Dmitri Mendeleev and the periodic table with the elements known in his time and the empty slots for elements still to be discovered. *Source:* Wikipedia, https://en.wikipedia.org/wiki/Dmitri_Mendeleev#/media/File:Dmitri_Mendeleev_1890s.jpg.

1.5 The Hydrogen Spectrum Lines

Johann Balmer (1825–1898), a Swiss mathematician, found empirically in 1895 that the separation of the optical lines generated by hydrogen gas can be expressed by a formula using just a constant, C , and integer numbers. He expressed his observation with Eq. (1.3)

$$\lambda = \frac{Cm^2}{m^2 - n^2} \quad (1.3)$$

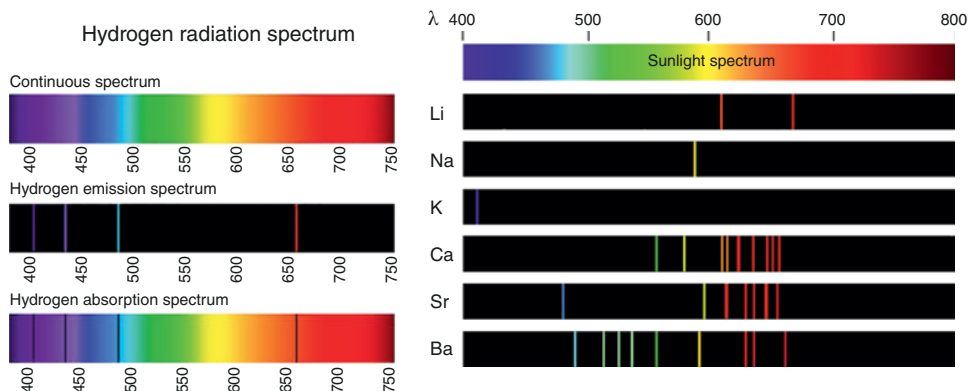


Figure 1.6 The spectrum of the hydrogen atom on the left shows the absorption lines (below) and the emission lines (middle). On the right are the emission lines of several other materials. *Source:* <https://www.shutterstock.com/image-vector/spectrum-spectral-line-example-hydrogen-emission-1288942888?src=iUiOwiDEznOcV6XzswXhMA-1-0> (left); <https://www.shutterstock.com/image-vector/line-spectra-elements-339037577?src=l6tWF1qlh6XcWayXsZl-Gw-3-16> (right). (See insert for a color representation of this figure.)

here λ is the wavelength of the missing line, C is a heuristically obtained constant ($C = 3.64 \times 10^{-9}$ m), $n = 2$, and m is an integer greater than 2 (i.e. 3, 4, 5, and so on). When you put any of these integer numbers in Eq. (1.3), you get the wavelength of all the lines in the hydrogen spectrum.

Figure 1.6 shows the hydrogen spectrum on the left, with its characteristic emission and absorption lines. These are the lines that Balmer used to develop Eq. (1.3) to calculate the missing hydrogen's wavelengths. All the elements have similar absorption and emission lines at different wavelengths, and I show a few on the right in Figure 1.6.

Just three years later, Johannes Rydberg (1854–1919) found that the Balmer equation was one specific case of a more general formula, Eq. (1.4):

$$\nu = \frac{1}{\lambda} = R \left[\frac{1}{n^2} - \frac{1}{m^2} \right] \quad (1.4)$$

The reciprocal of the wavelength is now given by a constant R and the same integer numbers, except that now n is allowed to have different integer numbers: 2, 3, 4, and so on. R is also a heuristically derived constant ($R = 1.1 \times 10^7 \text{ m}^{-1}$), called the *Rydberg constant*. Both Balmer and Rydberg (Figure 1.7) were able to quantify the entire spectrum of the hydrogen atom using the relationship in Eq. (1.4). It is interesting that Niels Bohr, whom I'll talk more about in Section 1.8, was able to calculate the Rydberg number using fundamental physical values, such as the mass of the electron, the electronic charge, the permittivity of free space, Planck's constant, and the speed of light (see Appendix 1.3). This behavior screams for an explanation.



Figure 1.7 Johann Balmer (left) found a mathematical relation for hydrogen’s spectral lines, and Johannes Rydberg (right) came up with a more general equation applicable to all gases and materials. *Source* Wikipedia, https://en.wikipedia.org/wiki/Johann_Jakob_Balmer#/media/File:Balmer.jpeg (left); Wikipedia, [https://en.wikipedia.org/wiki/Johannes_Rydberg#/media/File:Rydberg,_Janne_\(foto_Per_Bagge;_AFs_Arkiv\).jpg](https://en.wikipedia.org/wiki/Johannes_Rydberg#/media/File:Rydberg,_Janne_(foto_Per_Bagge;_AFs_Arkiv).jpg) (right).

1.6 Light is a Particle

Albert Einstein (1879–1955, Figure 1.8) published a paper in 1905 on the theory of the photoelectric effect. When light strikes a metal surface, it frees an electron if its energy is higher than a given threshold value. Any remaining energy is used to kick the electron off the surface. In his paper, Einstein proposed the concept that light has a dual personality; it behaves like a wave or like a particle, and the particle has an energy associated with the wavelength of that light.

He called this particle a “light quantum.” (In 1926, a French physicist named Frithiof Wolfers [1891–1971] renamed the light quantum a *photon*. It is interesting that Einstein received the Nobel Prize in 1921 for the discovery of the photon, not for his much more famous work on relativity.) This light particle, the photon, has an energy that depends on the frequency of the light. The energy associated with this light is given by the formula

$$E = hf = \frac{hc}{\lambda} \quad (1.5)$$

where h is Planck’s constant ($h = 6.63 \times 10^{-34} \text{ m}^2 \text{ kg s}^{-1}$), c is the speed of light ($c = 3 \times 10^8 \text{ m s}^{-1}$), and λ is the wavelength (m). The meter in the numerator cancels the one in the denominator, resulting in the energy given in Joules ($= \text{kg m}^2 \text{ s}^{-2}$).



Figure 1.8 Around 1905, Albert Einstein came up with the concept that light behaves as both a wave and a particle. *Source:* Wikipedia, https://en.wikipedia.org/wiki/Albert_Einstein#/media/File:Einstein_patentoffice.jpg.

1.7 The Atom's Structure

While all of these light experiments and relationships were being observed in the late nineteenth century, other scientists were playing with cathode-ray tubes, the precursors of old television sets and oscilloscopes, trying to understand the nature of the atom. The cathode-ray tube consists of an evacuated tube with two contacts, one at each end: the *cathode* and the *anode*. When a voltage is applied across the tube, current flows from the cathode to the anode, and the tube glows. The scientists explained this phenomenon by saying that electrons going through an evacuated tube containing very few atoms are able to attain sufficient velocity (and therefore kinetic energy) to hit the atoms and make them glow. They were called *cathode rays*.

Nobel Prize winning British physicist Joseph John Thomson (1856–1940, Figure 1.9) studied cathode rays and postulated in 1897 that they consisted of extremely small negatively charged particles, which he initially called “corpuscles.” (As happened with the term *photon*, George Stoney (1826–1911) later renamed corpuscles as *electrons*.) By studying how these particles moved through the gas and how they could be deflected by magnets, Thomson concluded that the “corpuscles” were (i) negatively charged particles and (ii) much smaller than the atoms themselves – at least 1000 times smaller. To account for electrically neutral atoms, he proposed that there is a core of positive charges with a large mass surrounded by an amorphous cloud of negatively charged electrons.

Ernest Rutherford (1871–1937, Figure 1.10), also a Nobel Prize winner, worked with radioactivity. In 1911, he bombarded very thin gold foil with alpha particles and looked at the scattered reflections as the radiation went through the foil. Most of the radiation went through the foil undeflected. Only a few alpha particles were reflected back and, from the angle of the reflected radiation, he concluded that the atom must have a very small,

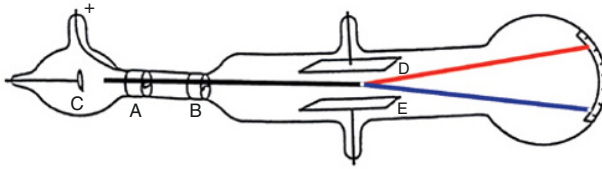


Figure 1.9 Joseph John Thomson and his cathode ray tube. *Source:* Wikipedia, https://en.wikipedia.org/wiki/J._J._Thomson#/media/File:JJ_Thomson_Cathode_Ray_2.png (left); Wikipedia, https://en.wikipedia.org/wiki/J._J._Thomson#/media/File:JJ_Thomson.jpg (right).

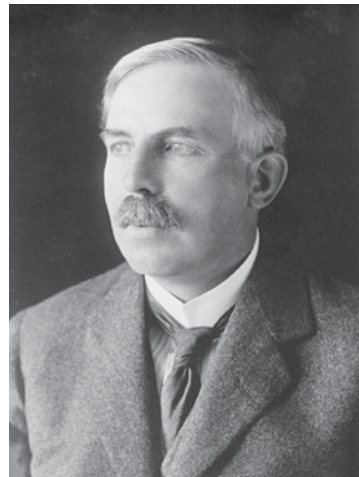
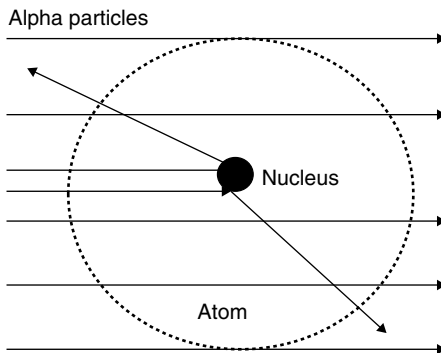


Figure 1.10 Ernest Rutherford, with his experiment that bombarded alpha particles with radiation, concluded that the nucleus is extremely small and is concentrated at the center of the atom. *Source:* Wikipedia, https://upload.wikimedia.org/wikipedia/commons/6/6e/Ernest_Rutherford_LOC.jpg.

concentrated, positively charged core to compensate for the negatively charged electrons. Because the large majority of alpha particles passed through the foil without any directional change, he concluded that the majority of the space in an atom is empty, and the electrons are orbiting the nucleus instead of just being a scrambled negatively charged cloud as Thomson had suggested.

Robert Millikan (1868–1953, Figure 1.11) was able to measure the electrical charge of an electron with an interesting oil drop experiment in 1909. He suspended a very small charged

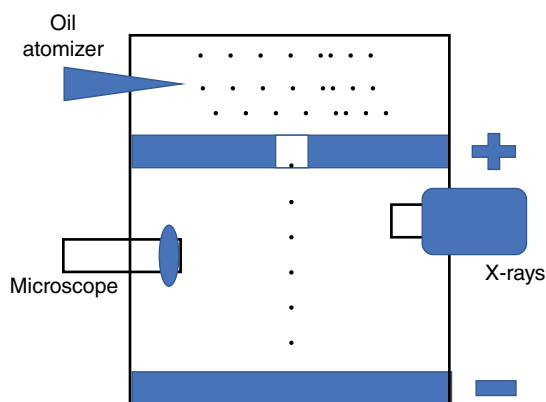


Figure 1.11 Robert Millikan, with his oil-drop experiment, measured the electrical charge of an electron. *Source:* https://en.wikipedia.org/wiki/Robert_Andrews_Millikan#/media/File:Millikan.jpg.

oil drop between two metal plates – one positive and the other negative – creating an electric field between the plates. He dropped tiny oil droplets into a vacuum chamber, and with X-rays, he negatively charged some of the oil drops. By changing the electric field between the two plates, he could control the speed of the oil drops, slowing them down, stopping them, or even moving them upward. By knowing the density of the oil drop, the size of the drop, its volume and mass, and the electric field that compensated for the effect of gravity, he was able to come up with the value of the charge of a single electron: 1.592×10^{-19} coulombs (he was off by less than 1% of the now-established number – not bad at all).

1.8 The Bohr Atom

So here we are in 1913 (just a mere 105 years ago at the time of this writing). What did Bohr know? He knew:

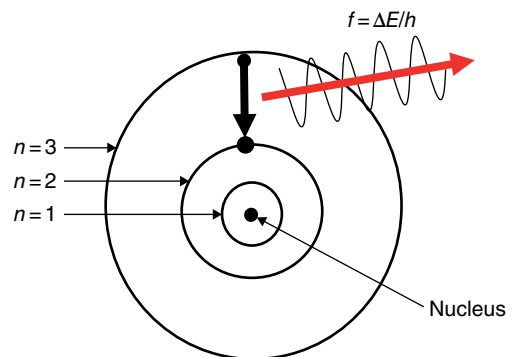
- 1) That a hydrogen atom is the simplest atom, consisting of just one proton (positively charged) and one electron (negatively charged).
- 2) That all of the atom's mass is concentrated at the core: that is, the proton.
- 3) That electrons are negative particles somehow orbiting the nucleus.
- 4) That the great majority of space in an atom is empty.
- 5) That all the other elements can be organized neatly by weight on a periodic table.
- 6) That all elements have different emission spectra with specific emission or absorption color lines.

Niels Bohr (1885–1962, Figure 1.12) was able to beautifully explain all of these observations and how the spectral lines are generated. He postulated in 1912 that an atom consists of a core nucleus that has all the mass and is surrounded by electrons, moving like a planetary system in well-defined orbits (Figure 1.13). Electrostatic forces between the proton



Figure 1.12 Niels Bohr (left) postulated the planetary model of the atom. Wolfgang Pauli (right), using quantum mechanics, proved that no two electrons in a system can have the same quantum numbers. *Source:* Wikipedia, https://en.wikipedia.org/wiki/Niels_Bohr#/media/File:Niels_Bohr.jpg (left); Wikipedia, https://en.wikipedia.org/wiki/Wolfgang_Pauli#/media/File:Pauli.jpg (right).

Figure 1.13 The Bohr planetary model of an atom has discrete and stable orbits. An electron falling from level 3 to level 2 transfers its energy to an equivalently energetic photon.



and the electron (analogous to the gravitational forces in the solar system) keep the electrons circulating without escaping their orbit. Additionally, Bohr postulated that the electrons in orbit do not radiate any energy, so the orbits are stable. The only way to radiate or absorb any energy is for an electron to jump from one orbit to another, and that is precisely what explains the spectra of hydrogen and other elements.

Since electrons are forbidden to have any energy except for the energy of a specific orbit, they have to jump from one orbit to another, like going up the stairs, one, two, or three steps at a time (not one and a half). When falling from a higher orbit to a lower one, the electron releases a fixed packet of energy in the form of a photon of a very precise frequency (remember that Einstein said light behaves like a particle with an energy related to the wavelength of the light: Eq. 1.5). The transition from orbit 3 to orbit 2, as I show in

Figure 1.13, results in the emission of a photon of a very precise frequency, given by the change in energy, ΔE , divided by Planck's constant. Similarly, if an electron in orbit 2 wants to jump to orbit 3, the hydrogen atom has to absorb the energy it needs by absorbing a photon with the same precise energy, or by thermal heating, or by some other means. All other light photons not exactly matched to the difference between the energy levels go through the material unimpaird. The material is therefore transparent for all of the light waves that do not match the exact difference between two energy levels.

In 1924, Austrian Wolfgang Pauli (1900–1958, on the right in Figure 1.12) proposed his exclusion principle, which states that no two electrons (or fermion particles) in a system can have the same quantum numbers. The first atomic level of any element can hold only 2 electrons, the second 8, the third 18, the fourth 32, etc. A simple relation tells you how many electrons can share a given energy orbit: $2n^2$. You may wonder why. If, according to Pauli's exclusion principle, the electrons cannot share the same quantum state, why do we have more than one electron in each orbit? The answer is that each electron is described by four quantum numbers (like the three numbers that describe your first, middle, last names, and your date of birth), but only the first quantum number, n , specifies the energy of the electron and thus explains the behavior of the light spectra. I explain the electron's four quantum numbers in more detail in Appendix 1.1.

Here's an analogy. Suppose I have a theater with 2 seats in the first row, 8 in the second, 18 in the third, 32 in the fourth, etc. The tickets for the first row cost \$20, the second row \$50, the third row \$75, the fourth \$125, and so on. (I know, it is a weird theater, but this is just an analogy.) Spectators are forbidden to sit in someone else's lap or stand in the aisles. If 12 people show up for the performance, they first occupy the 2 seats of the first row, the next 8 patrons occupy the second row, and the last 2 spectators sit somewhere in the third row. Further back in the theater, the rows have more seats, but they are empty. If a patron wants to change rows – from row 3 to row 4, for example – he has to pay the extra \$50: the difference in the price of the tickets in the different rows. If he moves the other way, from row 4 to row 3, he is reimbursed the \$50. Now, if a wealthy person in row 1 wants to move to row 4, he will be required to pay \$105. That is, money must be paid or received to move from one row to another. All these changes assume that the seat someone desires is unoccupied. If the group of spectators is short of money (no energy), they will occupy the seats of the first rows as long as there are seats available. If the group is wealthy (has lots of energy), they can jump from one seat to another as long as they have enough money (energy) to afford the higher prices. The amount they have to pay depends only on the difference in the price of the seats in each row. End of analogy.

At 0K, absolute temperature (-273°C), there is no energy whatsoever, so all the electrons occupy the lowest allowed energy levels. At room temperature, 300°C , there is quite a large amount of thermal energy, and electrons start moving from one level to another, leaving empty seats that can be occupied by other electrons, absorbing or emitting photons as they move.

Figure 1.14 shows the transitions observed in the hydrogen atom. The groups of lines were named later by those who found them.

Have you ever wondered why, when we walk on the second floor, we do not fall through it and land on the first floor? Think about it. The typical size of an atom is 5×10^{-10} m, and the size of a nucleus is about 30000 times smaller, 1.6×10^{-15} m. All the mass is