Introduction to Semiconductors

CHAPTER OUTLINE

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INTROdUCTION

Electronic devices such as diodes, transistors, and integrated circuits are made of a semi-conductive material. To understand how these devices work, you should have basic knowledge of the structure of atoms and interaction of the atomic particles. An important concept introduced in this chapter is that of the *p-n* junction that is formed when two different types of semi-conductive material are joined. The *p-n* junction is fundamental to the operation of devices such as the solar cell, the diode, and certain types of transistors.

CHAPTER OBJECTIVES

- Describe the structure of an atom
- Discuss insulators, conductors, and semiconductors and how they differ
- Describe how current is produced in a semiconductor
- Describe the properties of *n*-type and *p*-type semiconductors
- Describe how a *p-n* junction is formed

KEY TERMS

- Atom
- Proton
- Electron
- Shell
- Valence
- Ionization
- Free electron
- Orbital
- Insulator
- Conductor
- Semiconductor
- Silicon
- Crystal
- Hole
- Metallic bond
- Doping
- *PN* junction
- Barrier potential

1–1 The Atom

All matter is composed of atoms; all atoms consist of electrons, protons, and neutrons except normal hydrogen, which does not have a neutron. Each element in the periodic table has a unique atomic structure, and all atoms for a given element have the same number of protons. At first, the atom was thought to be a tiny indivisible sphere. Later it was shown that the atom was not a single particle but was made up of a small, dense nucleus around which electrons orbit at great distances from the nucleus, similar to the way of planets orbit of the Sun. Niels Bohr proposed that the electrons in an atom circle the nucleus in different obits, similar to the way planets orbit around the sun in our solar system. The Bohr model is often referred to as the planetary model. Another view of the atom called the *quantum model* is considered a more accurate representation, but it is difficult to visualize. For most practical purposes in electronics, the Bohr model suffices and is commonly used because it is easy to visualize.

After completing this section, you should be able to

- **Describe the structure of an atom**
  - Discuss the Bohr model of an atom
    - Define *electron*, *proton*, *neutron*, and *nucleus*
    - Define *atomic number*
    - Discuss electron shells and orbits
    - Explain energy levels
    - Define *valence electron*
    - Discuss ionization
      - Define *free electron* and *ion*
    - Discuss the basic concept of the quantum model of the atom

HISTORY

Niels Henrik David Bohr (October 7,1885–November 18, 1962) was a Danish physicist, who made important contributions to understanding the structure of the atom and quantum mechanics by postulating the "planetary" model of the atom. He received the Nobel Prize in physics in 1922. Bohr drew upon the work or collaborated with scientists such as Dalton, Thomson, and Rutherford, among others and has been described as one of the most influential physicists of the 20th century.

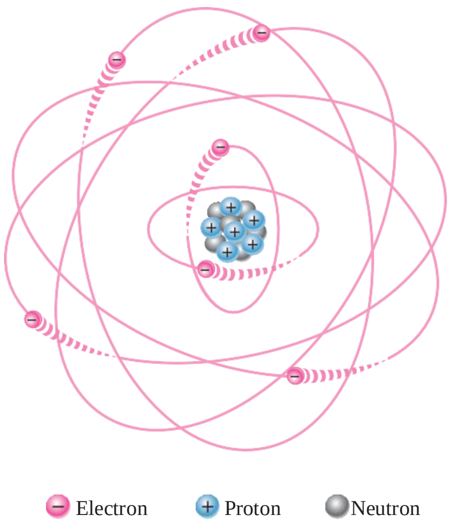
The Bohr Model

An **atom*** is the smallest particle of an element that retains the characteristics of that element. Each of the known 118 elements has atoms that are different from the atoms of all other elements. This gives each element a unique atomic structure. According to the classical Bohr model, atoms have a planetary type of structure that consists of a central nucleus surrounded by orbiting electrons, as illustrated in Figure 1–1. The **nucleus** consists of positively charged particles called **protons** and uncharged particles called **neutrons.** The basic particles of negative charge are called **electrons** that move around the nucleus.

Each type of atom has a certain number of electrons and protons that distinguishes it from the atoms of all other elements. For example, the simplest atom is that of hydrogen, which has one proton and one electron, as shown in Figure 1–2(a). As another example, the helium atom, shown in Figure 1–2(b), has two protons and two neutrons in the nucleus and two electrons orbiting the nucleus.

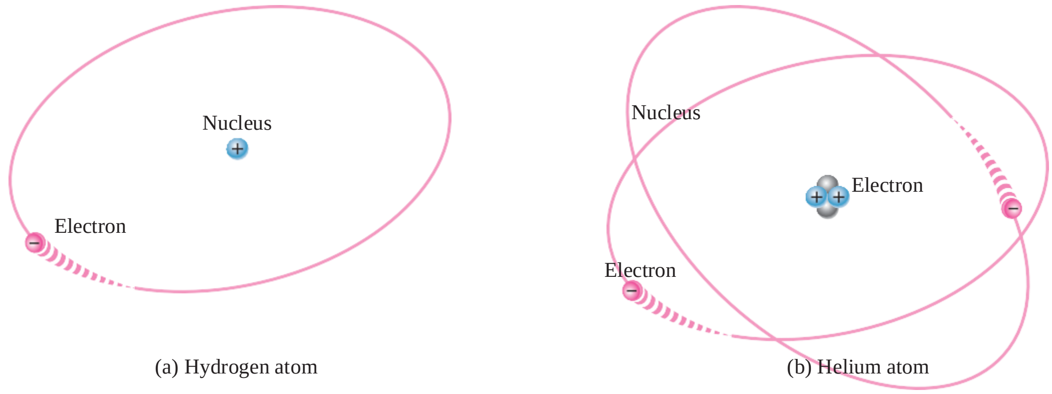
Atomic Number

All elements are arranged in the periodic table of the elements in order according to their atomic number. The **atomic number** equals the number of protons in the nucleus, which is the same as the number of electrons in an electrically balanced (neutral) atom. For example, hydrogen has an atomic number of 1 and helium has an atomic number of 2. In their normal (or neutral) state, all atoms of a given element have the same number of electrons as protons; the positive charges cancel the negative charges, and the atom has a net charge of zero.



▲ FIGURE 1–1

The Bohr model of an atom showing electrons in orbits around the nucleus, which consists of Protons and Neutrons. The "tails" on the Electrons indicate motion.



▲ **FIGURE 1–2** Two simple atoms, Hydrogen and Helium.

Atomic numbers of all the elements are shown on the periodic table of the elements in Figure 1–3 below.

Helium
Atomic number = 2

Silicon
Atomic number = 14

1 **H**																	2 **He**
3 **Li**	4 **Be**											5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
11 **Na**	12 **Mg**											13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**
19 **K**	20 **Ca**	21 **Sc**	22 **Ti**	23 **V**	24 **Cr**	25 **Mn**	26 **Fe**	27 **Co**	28 **Ni**	29 **Cu**	30 **Zn**	31 **Ga**	32 **Ge**	33 **As**	34 **Se**	35 **Br**	36 **Kr**
37 **Rb**	38 **Sr**	39 **Y**	40 **Zr**	41 **Nb**	42 **Mo**	43 **Tc**	44 **Ru**	45 **Rh**	46 **Pd**	47 **Ag**	48 **Cd**	49 **In**	50 **Sn**	51 **Sb**	52 **Te**	53 **I**	54 **Xe**
55 **Cs**	56 **Ba**	*	72 **Hf**	73 **Ta**	74 **W**	75 **Re**	76 **Os**	77 **Ir**	78 **Pt**	79 **Au**	80 **Hg**	81 **Tl**	82 **Pb**	83 **Bi**	84 **Po**	85 **At**	86 **Rn**
87 **Fr**	88 **Ra**	**	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**	110 **Ds**	111 **Rg**	112 **Cp**	113 **Uut**	114 **Uuq**	115 **Uup**	116 **Uuh**	117 **Uus**	118 **Uuo**

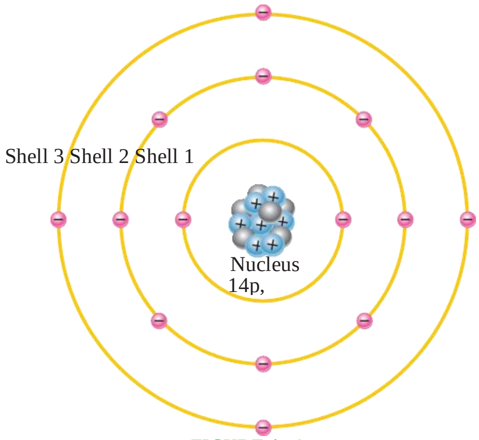
57 **La**	58 **Ce**	59 **Pr**	60 **Nd**	61 **Pm**	62 **Sm**	63 **Eu**	64 **Gd**	65 **Tb**	66 **Dy**	67 **Ho**	68 **Er**	69 **Tm**	70 **Yb**	71 **Lu**
89 **Ac**	90 **Th**	91 **Pa**	92 **U**	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

▲ **FIGURE 1–3**
The periodic Table of the elements. Some tables also show atomic mass.

Electrons and Shells

Energy Levels:- Electrons orbit around the nucleus of an atom at certain distances from the nucleus. Electrons near the nucleus have less energy than those in more distant orbits. Only discrete (separate and distinct) values of electron energies exist within atomic structures. Therefore, electrons must orbit only at discrete distances from the nucleus.

Each discrete distance (**orbit**) from the nucleus corresponds to a certain energy level. In an atom, the orbits are grouped into energy levels known as **shells**. A given atom has a fixed number of shells. Each shell has a fixed maximum number of electrons. The shells (energy levels) are designated 1, 2, 3, and so on, with 1 being closest to the nucleus. The Bohr model of the silicon atom is shown in Figure 1–4. Notice that there are 14 electrons surrounding the nucleus with exactly 14 protons, and usually 14 neutrons in the nucleus.



▸ FIGURE 1–4
Illustration of the Bohr model of the silicon atom.

The Maximum Number of Electrons in Each Shell

The maximum number of electrons (N_e) that can exist in each shell of an atom is a fact of nature and can be calculated by the formula,

Equation 1–1 $$N_e = 2n^2$$

where n is the number of the shell. The maximum number of electrons that can exist in the innermost shell (shell 1) is

$$N_e = 2n^2 = 2(1)^2 = 2$$

The maximum number of electrons that can exist in shell 2 is

$$N_e = 2n^2 = 2(2)^2 = 2(4) = 8$$

The maximum number of electrons that can exist in shell 3 is

$$N_e = 2n^2 = 2(3)^2 = 2(9) = 18$$

The maximum number of electrons that can exist in shell 4 is

$$N_e = 2n^2 = 2(4)^2 = 2(16) = 32$$

Atoms are extremely small and cannot be seen even with the strongest optical microscopes; however, a scanning tunneling microscope can detect a single atom. The nucleus is so small and the electrons orbit at such distances that the atom is mostly empty space. To put it in perspective, if the proton in a hydrogen atom were the size of a golf ball, the electron orbit would be approximately one mile away.

Protons and Neutrons are approximately the same mass. The mass of an electron is 1/1836 of a proton. Within Protons and Neutrons, there are even smaller particles called quarks. Quarks are the subject of intense study by particle, physicists as they explain the existence of more than 100 subatomic particles.

Valence Electrons

Electrons that are in orbits farther from the nucleus have higher energy and are less tightly bound to the atom than those closer to the nucleus. This is because the force of attraction between the positively charged nucleus and the negatively charged electron decreases with increasing distance from the nucleus. Electrons with the highest energy exist in the outermost shell of an atom and are relatively loosely bound to the atom. This outermost shell is known as the **valence** shell, and electrons in this shell are called *valence electrons*. These valence electrons contribute to chemical reactions and bonding within the structure of a material and determine its electrical properties. When a valence electron gains sufficient energy from an external source, it can break free from its atom. This is the basis for conduction in materials.

Ionization

When an atom absorbs energy, the valence electrons can easily jump to higher energy shells. If a valence electron acquires a sufficient amount of energy, called *ionization energy*, it can actually escape from the outer shell and the atom's influence. The departure of a valence electron leaves a previously neutral atom with an excess of positive charge (more protons than electrons). The process of losing a valence electron is known as **ionization**, and the resulting positively charged atom is called a *positive ion*. For example, the chemical symbol for hydrogen is H. When a neutral hydrogen atom loses its valence electron and becomes a positive ion, it is designated H^+. The escaped valence electron is called a **free electron**.

The reverse process can occur in certain atoms when a free electron collides with the atom and is captured, releasing energy. The atom that has acquired the extra electron is called a *negative ion*. The ionization process is not restricted to single atoms. In many chemical reactions, a group of atoms that are bonded together can lose or acquire one or more electrons.

For some nonmetallic materials such as chlorine, a free electron can be captured by the neutral atom, forming a negative ion. In the case of chlorine, the ion is more stable than the neutral atom because it has a filled outer shell. The chlorine ion is designated as Cl^-.

The Quantum Model

Although the Bohr model of an atom is widely used because of its simplicity and ease of visualization, it is not a complete model. The quantum model is considered to be more accurate. The quantum model is a statistical model and very difficult to understand or visualize. Like the Bohr model, the quantum model has a nucleus of protons and neutrons surrounded by electrons. Unlike the Bohr model, the electrons in the quantum model do not exist in precise circular orbits as particles. Three important principles underlie the quantum model: the **wave-particle duality principle**, the **uncertainty principle**, and the **superposition principle**.

Wave-particle duality	***Uncertainly principle***	***Superposition.***
Just as light can be thought of as exhibiting both a wave and a particle **(photon),** electrons are thought to exhibit a wave-particle duality. The velocity of an orbiting electron is related to its wavelength, which interferes with neighboring electron wavelengths by amplifying or canceling each other.	As you know, a wave is characterized by peaks and valleys; therefore, electrons acting as waves cannot be precisely identified in terms of their position. According to a principle ascribed to Heisenberg, it is impossible to determine simultaneously both the position and velocity of an electron with any degree of accuracy or certainty. The result of this principle produces a concept of the atom with **probability clouds**, which are mathematical descriptions of where electrons in an atom are most likely to be located.	A principle of quantum theory that describes a challenging concept about the behavior of matter and forces at the subatomic level. Basically, the principle states that although the state of any object is unknown, it is actually in all possible states simultaneously as long as an observation is not attempted. An analogy known as Schrodinger's cat is often used to illustrate in an oversimplified way quantum superposition.

NOTATION	EXPLANATION
$1s^2$	**2 electrons in shell 1, orbital *s***
$2s^2\ 2p^3$	**5 electrons in shell 2: 2 in orbital *s*, 3 in orbital *p***

Table 1-1
Electron configuration table for Nitrogen.

In the quantum model, each shell or energy level consists of up to four subshells called **orbitals**, which are designated *s*, *p*, *d*, and *f*. Orbital *s* can hold a maximum of two electrons, orbital *p* can hold six electrons, orbital *d* can hold 10 electrons, and orbital *f* can hold 14 electrons. Each atom can be described by an electron configuration table that shows the shells or energy levels, the orbital, and the number of electrons in each orbital. For example, the electron configuration table for the nitrogen atom is given in Table 1–1. The first full-size number is the shell or energy level, the letter is the orbital, and the exponent is the number of electrons in the orbital.

Atomic orbitals do not resemble a discrete circular path for the electron as depicted in Bohr's planetary model. In the quantum picture, each shell in the Bohr model is a three dimensional space surrounding the atom that represents the mean (average) energy of the electron cloud. The term **electron cloud** (probability cloud) is used to describe the area around an atom's nucleus where an electron will probably be found.

EXAMPLE 1–1 Using the atomic number from the periodic table in Figure 1–3, describe a silicon (Si) atom using an electron configuration table.

Solution The atomic number of silicon is 14. This means that there are 14 protons in the nucleus. Since there is always the same number of electrons as protons in a neutral atom, there are also 14 electrons. As you know, there can be up to two electrons in shell 1, eight in shell 2, and eighteen in shell 3. Therefore, in silicon there are two electrons in shell 1, eight electrons in shell 2, and four electrons in shell 3 for a total of 14 electrons. The electron configuration table for silicon is shown in Table 1–2.

▸ **TABLE 1–2**

NOTATION	EXPLANATION
$1s^2$	2 electrons in shell 1, orbital *s*
$2s^2$ $2p^6$	8 electrons in shell 2: 2 in orbital *s*, 6 in orbital *p*
$3s^2$ $3p^2$	4 electrons in shell 3: 2 in orbital *s*, 2 in orbital *p*

*Related Problem** Develop an electron configuration table for the germanium (Ge) atom in the periodic table.

*** Answers can be found at www.pearsonhighered.com/floyd.**

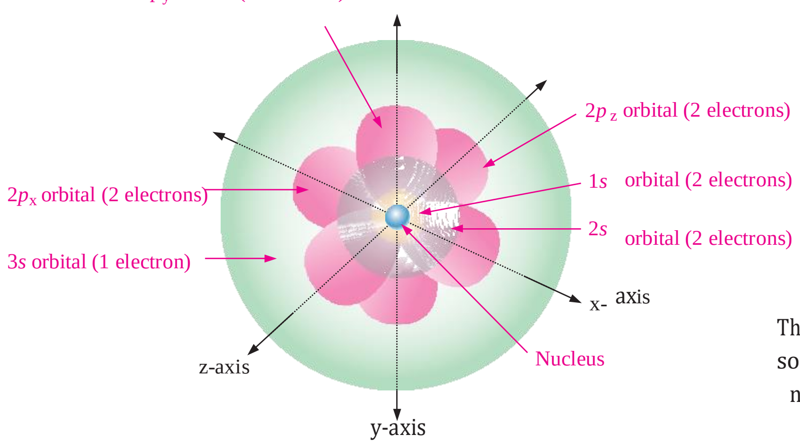


FIGURE 1–5
Three-dimensional quantum model of the sodium atom, showing the orbitals and number of electrons in each orbital.

In a three-dimensional representation of the quantum model of an atom, the *s*-orbitals are shaped like spheres with the nucleus in the center. For energy level 1, the sphere is a single sphere, but for energy levels 2 or more, each single *s*-orbital is composed of nested spherical shells. A *p*-orbital for shell 2 has the form of two ellipsoidal lobes with a point of tangency at the nucleus (sometimes referred to as a dumbbell shape.) The three *p*-orbitals in each energy level are oriented at right angles to each other. One is oriented on the *x*-axis, one on the *y*-axis, and one on the *z*-axis. For example, a view of the quantum model of a sodium atom (Na) that has 11 electrons as shown in Figure 1–5. The three axes are shown to give you a 3-D perspective.

Section 1-1 Check up

1. Describe the Bohr model of the atom. Define free *electron.*
2. What is the nucleus of an atom composed of?
3. Define each component. Define *atomic number.*
4. Discuss the difference between positive and negative ionization.
5. Discuss electron shells and orbits and their energy levels.
6. What is a valence electron?
7. Name three principles that distinguish the quantum model.
8. Because of its barrier potential, can a diode be used as a voltage source? Explain

*Answers can be found at www .pearsonhighered.com/floyd