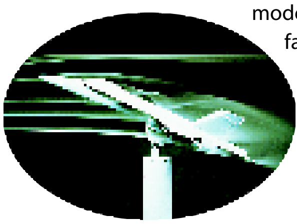
5.1 Models of the Atom

Connecting to Your World Aeronautical engineers use wind tunnels and scale models to simulate and test the forces from the moving air on each proposed design. The scale model shown is a physical model. However, not all models are physical. In fact, several theoretical models of the atom have been developed over the last few hundred years. In this section, you will learn about the currently accepted model of how electrons behave in atoms.



Guide for Reading

Key Concepts

- What was inadequate about Rutherford's atomic model?
- What was the new proposal in the Bohr model of the atom?
- What does the quantum mechanical model determine about the electrons in an atom?
- How do sublevels of principal energy levels differ?

Vocabulary

energy levels
quantum
quantum mechanical model
atomic orbital

Reading Strategy

Using Prior Knowledge Before you read, jot down three things you already know about atoms. When you read the section, explain how what you already knew helped you learn something new.

The Development of Atomic Models

So far in this textbook, the model for the atom consisted of protons and neutrons making up a nucleus surrounded by electrons. After discovering the atomic nucleus, Rutherford used existing ideas about the atom and proposed an atomic model in which the electrons move around the nucleus, like the planets move around the sun. Rutherford's model explained only a few simple properties of atoms. It could not explain, for example, why metals or compounds of metals give off characteristic colors when heated in a flame, or why objects—when heated to higher and higher temperatures—first glow dull red, then yellow, then white, as shown in Figure 5.1. **Rutherford's atomic model could not explain the chemical properties of elements.** Explaining what leads to the chemical properties of elements requires a model that better describes the behavior of electrons within atoms.

Figure 5.1 Rutherford's model fails to explain why objects change color when heated. As the temperature of this horseshoe is increased, it first appears black, then red, then yellow, and then white. The observed behavior could be explained only if the atoms in the iron gave off light in specific amounts of energy. A better atomic model was needed to explain this observation.

Section Resources

Print

- **Guided Reading and Study Workbook,** Section 5.1
- **Core Teaching Resources,** Section 5.1 Review
- **Transparencies,** T57

Technology

- **Interactive Textbook with ChemASAP,** Animation 5, Assessment 5.1

5.1

1 FOCUS

Objectives

5.1.1 **Identify** the inadequacies in the Rutherford atomic model.

5.1.2 **Identify** the new proposal in the Bohr model of the atom.

5.1.3 **Describe** the energies and positions of electrons according to the quantum mechanical model.

5.1.4 **Describe** how the shapes of orbitals related to different sublevels differ.

Guide for Reading

Build Vocabulary L2

Word Parts The word *orbital* was recently (1932) minted by scientists to describe the space in which an atom has a high probability of being found. It is derived from the word *orbit*, the path around the nucleus that electrons had been postulated to follow.

Reading Strategy L2

Relate Text and Visuals Have students follow the Time Line in Figure 5.2 and explain why each model was inadequate or incomplete.

2 INSTRUCT

Connecting to Your World

Have students study the photograph and read the text. Ask, **What are two types of models?** *(physical and theoretical)* **Why are models important?** *(You can test a model to see if your ideas or design are correct.)*

The Development of Atomic Models

Relate L2

Compare the evolution of the car to the development of the atomic model, in particular the role of scientific discoveries (semiconductors or subatomic particles) and new technology (assembly lines or particle accelerators).

The Bohr Model

Discuss L2

Prepare students for the expanded view of the atom by asking, **What are the three major subatomic particles that comprise an atom?** *(electron, proton, and neutron)* **What are the electrical charges associated with each of these particles?** *(electron, 1–; proton, 1+; neutron, 0)* **Describe the structure of the nuclear atom in terms of the locations of each of the subatomic particles.** *(Every atom, except hydrogen, consists of a small dense nucleus composed of protons and neutrons that accounts for most of the mass of the atom. Hydrogen is an exception because it has only one proton in its nucleus. Negatively charged electrons surround the nucleus and occupy most of its volume. Electrons contribute little to the mass of an atom.)*

Word *Origins* L2

The word *quantity* has the same root as *quantum* and means "an indefinite amount or number."

TEACHER Demo

Quantized Energy L2

Purpose Students witness an example of quantized energy.

Materials trumpet or trombone and a player

Procedure Have a student blow into a trumpet or trombone. Challenge the student to produce as many different notes as he or she can without depressing any valves.

Expected Outcomes It is not possible to play the entire scale without changing the valve positions. In the open position, no matter how much energy is put in, the instrument will only accept certain specific amounts of energy and produce certain notes. The instrument accepts quantized packages of energy.

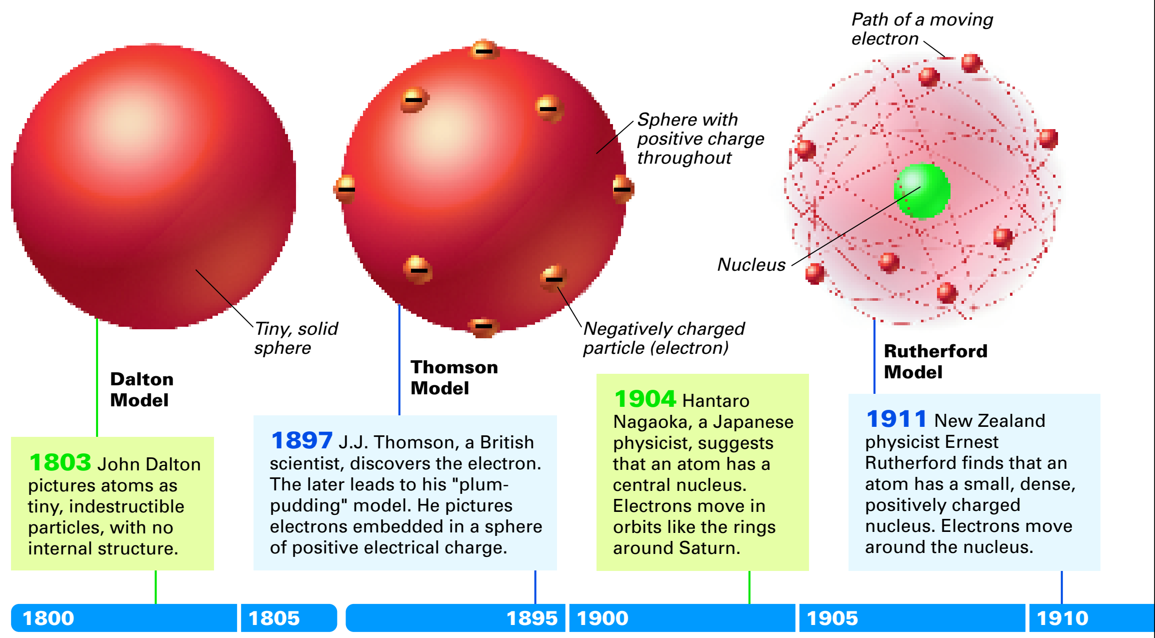


Figure 5.2 These illustrations show how the atomic model has changed as scientists learned more about the atom's structure.

The Bohr Model

Niels Bohr (1885–1962), a young Danish physicist and a student of Rutherford, believed Rutherford's model needed improvement. In 1913 Bohr changed Rutherford's model to include newer discoveries about how the energy of an atom changes when it absorbs or emits light. He considered the simplest atom, hydrogen, which has one electron. **Bohr proposed that an electron is found only in specific circular paths, or orbits, around the nucleus.** The timeline in Figure 5.2 shows the development of atomic models from 1800 to 1935.

Each possible electron orbit in Bohr's model has a fixed energy. The fixed energies an electron can have are called **energy levels.** The fixed energy levels of electrons are somewhat like the rungs of the ladder in Figure 5.3a. The lowest rung of the ladder corresponds to the lowest energy level. A person can climb up or down a ladder by going from rung to rung. Similarly, an electron can jump from one energy level to another. A person on a ladder cannot stand between the rungs. Similarly, the electrons in an atom cannot be between energy levels. To move from one rung to another, a person climbing a ladder must move just the right distance. To move from one energy level to another, an electron must gain or lose just the right amount of energy. In general, the higher an electron is on the energy ladder, the farther it is from the nucleus.

A **quantum** of energy is the amount of energy required to move an electron from one energy level to another energy level. The energy of an electron is said to be quantized. You have probably heard the term *quantum leap* used to describe an abrupt change. The term originates from the ideas found in the Bohr model of the atom.

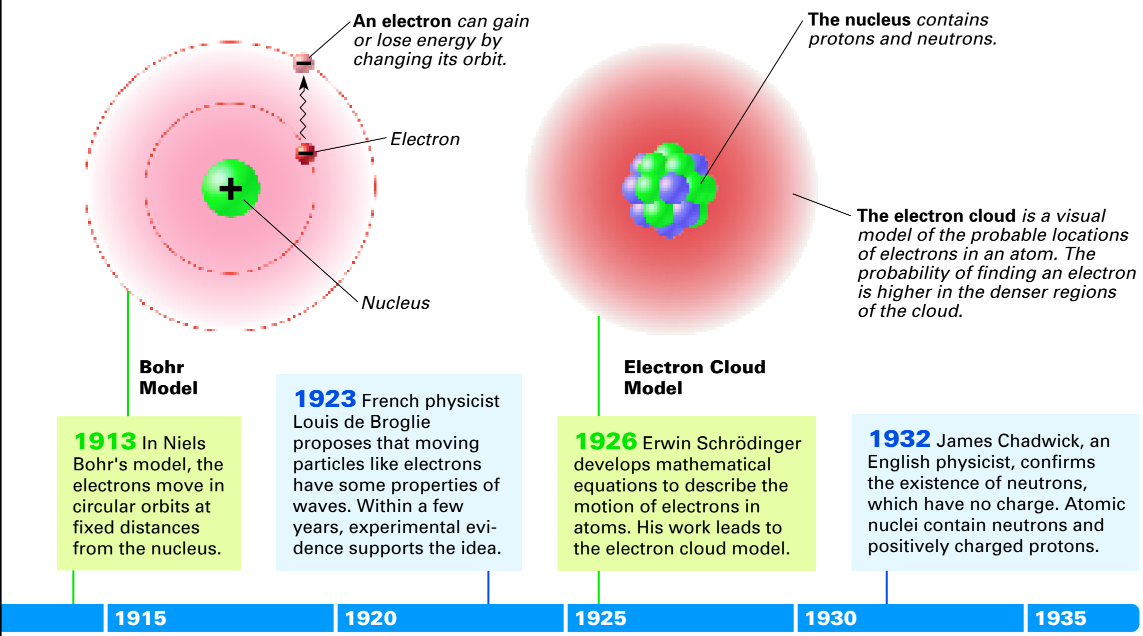
Word *Origins*

Quantum comes from the Latin word *quantus,* meaning "how much." **What other commonly used English word comes from this root?**

Differentiated Instruction

English Learners L2

Have students make a poster to show the evolution of the model of the atom. Have them start with Dalton's model and continue through all the modifications to the current quantum mechanical model. Have them include enough information to make the model clear and understandable.



The amount of energy an electron gains or loses in an atom is not always the same. Like the rungs of the strange ladder in Figure 5.3b, the energy levels in an atom are not equally spaced. The higher energy levels are closer together. It takes less energy to climb from one rung to another near the top of the ladder in Figure 5.3b, where the rungs are closer. Similarly, the higher the energy level occupied by an electron, the less energy it takes to move from that energy level to the next energy level.

The Bohr model gave results in agreement with experiment for the hydrogen atom. However, it still failed in many ways to explain the energies absorbed and emitted by atoms with more than one electron.

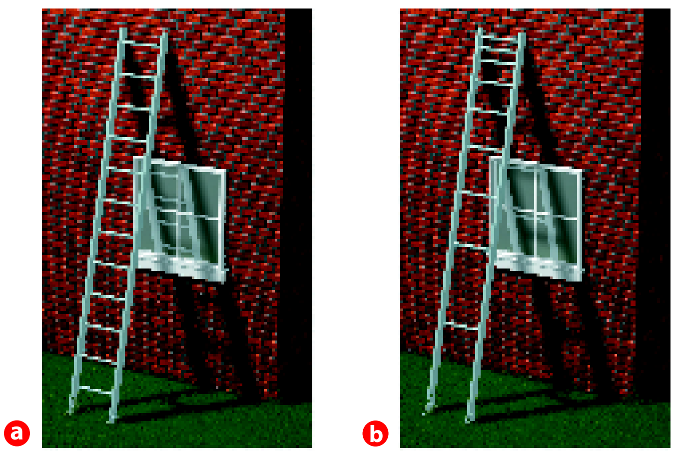


Figure 5.3 These ladder steps are somewhat like energy levels. **a** In an ordinary ladder, the rungs are equally spaced. **b** The energy levels in atoms are unequally spaced, like the rungs in this ladder. The higher energy levels are closer together.

Use Visuals L1

Figure 5.3 Ask, **What are the rungs in the ladders compared to?** *(the energy levels in atoms)* **How is the ladder in b different from the ladder in a?** *(The rungs are not evenly spaced. They get closer together nearer to the top.)*

Relate L2

When light shines on a fluorescent material, its electrons absorb the energy and move to a higher energy level. Almost immediately, the material begins to emit light as the electrons drop back down to their usual energy level. Have students bring objects painted with fluorescent paint or glow-in-the-dark objects to class and observe them in a darkened room.

CLASS Activity

Energy and Energy Levels L1

Purpose Students observe an analogy for energy transitions in atoms.

Materials one red, one yellow, and one blue ball, step stool with three steps

Procedure Stand on the first step of the step stool. Jump down to the floor while throwing the red ball to a student catcher. Stand on the second step and repeat the process throwing the yellow ball. Repeat jumping from the third step while throwing the blue ball. Ask students to state the analogy. *(The different colored balls are different amounts of energy that correspond to the energy emitted when electrons drop from excited energy levels to a lower level. Blue must be the largest amount of energy and red the smallest.)*

The Quantum Mechanical Model

Discuss

L2

Have students imagine a used dartboard that has concentric circles of decreasing value. (The analogy will not work for professional dartboards.) Ask students to describe the distribution of holes. *(More holes are close to the bull's-eye than near the edge, which is analogous to the distribution of electrons in an atom. Probability diminishes with distance from the nucleus, approaching zero as distance approaches infinity.)* Ask, **If the dartboard is a model for an atom, what are its major failures?** *(It is a 2-dimensional model for a 3-dimensional atom; the distribution of holes may not decrease uniformly)* Have students assume that the holes do decrease uniformly. Ask them to compare the probability for two holes on a circle at a specified distance from the center. *(The probabilities are equal.)* This rule holds true for electrons that are equidistant from the nucleus.

CLASS Activity

The Shapes of Orbitals

L2

Purpose Students observe the shapes of *s* and *p* orbitals and their orientation in a three-dimensional coordinate system.

Materials Three metal or wooden skewers, three Styrofoam balls (3 or 4 inches in diameter)

Procedure Push three skewers through a ball at right angles to each other and as close to the ball's center as possible. Ask students what type of orbital it represents. (*s*) Ask them to describe its shape. (*Students may say it is circular. Make the point that it is spherical.*) Identify the *x*-, *y*-, and *z*-axes. Push a skewer through two balls to represent a *p* orbital. Tell students that the skewer can represent any one of the three axes. Ask, **Where along the axis is the atomic nucleus?** *(at the point between the two balls)*

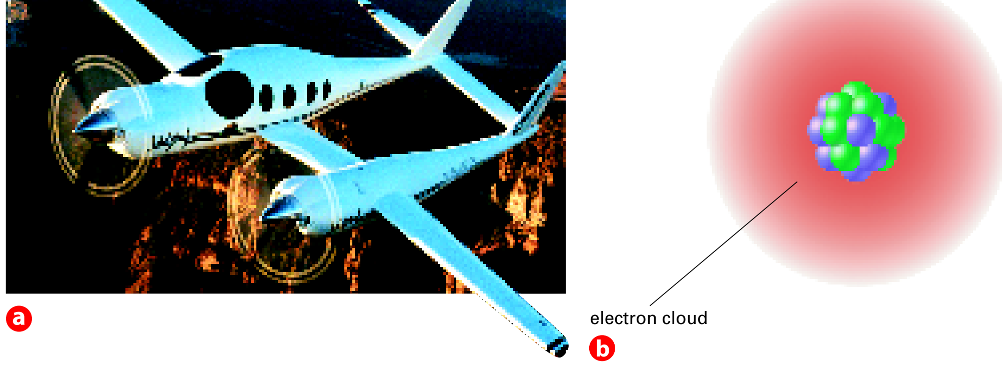


Figure 5.4 The electron cloud of an atom is compared here to photographs of a spinning airplane propeller. **a** The airplane propeller is somewhere in the blurry region it produces in this picture, but the picture does not tell you its exact position at any instant. **b** Similarly, the electron cloud of an atom represents the locations where an electron is likely to be found.

The Quantum Mechanical Model

The Rutherford planetary model and the Bohr model of the atom are based on describing paths of moving electrons as you would describe the path of a large moving object. New theoretical calculations and experimental results were inconsistent with describing electron motion this way. In 1926, the Austrian physicist Erwin Schrödinger (1887–1961) used these new results to devise and solve a mathematical equation describing the behavior of the electron in a hydrogen atom. The modern description of the electrons in atoms, the **quantum mechanical model,** comes from the mathematical solutions to the Schrödinger equation.

Like the Bohr model, the quantum mechanical model of the atom restricts the energy of electrons to certain values. Unlike the Bohr model, however, the quantum mechanical model does not involve an exact path the electron takes around the nucleus. **The quantum mechanical model determines the allowed energies an electron can have and how likely it is to find the electron in various locations around the nucleus.**

How likely it is to find the electron in a particular location is described by probability. If you place three red marbles and one green marble into a box and then pick a marble without looking, the probability of picking the green marble is one in four, or 25%. This means that if you put the four marbles in a box and picked one, and repeated this a great many times, you would pick a green marble in 25% of your tries.

The quantum mechanical model description of how the electron moving around the nucleus is similar to the motion of a rotating propeller blade. Figure 5.4a shows that the propeller blade has the same probability of being anywhere in the blurry region it produces in the picture, but you cannot tell its precise location at any instant. Similarly, in the quantum mechanical model of the atom, the probability of finding an electron within a certain volume of space surrounding the nucleus can be represented as a fuzzy cloud. The cloud is more dense where the probability of finding the electron is high. The cloud is less dense where the probability of finding the electron is low. Though it is unclear where the cloud ends, there is at least a slight chance of finding the electron at a considerable distance from the nucleus. Therefore, attempts to show probabilities as a fuzzy cloud are usually limited to the volume in which the electron is found 90% of the time. To visualize an electron probability cloud, imagine that you could mold a sack around the cloud so that the electron was inside the sack 90% of the time. The shape of the sack would then give you a useful picture of the shape of the cloud.

Facts and Figures

Development of Wave Mechanics

Schrödinger's equation accounted mathematically for de Broglie's discovery of the wave-like properties of the electron. These developments led to a major break away from the Bohr model of the atom.

Schrödinger extended de Broglie's work by considering the movement of a particle in an electromagnetic field, which established an area of physics known as wave mechanics.

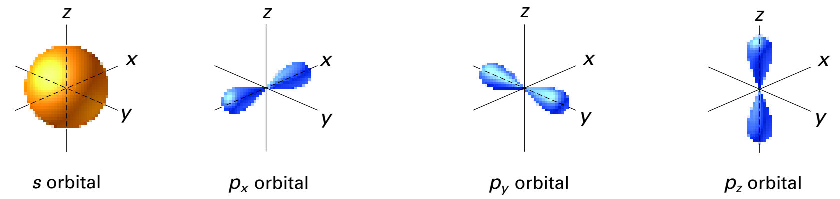


Figure 5.5 The electron clouds for the *s* orbital and the *p* orbitals are shown here.

Atomic Orbitals

Solving the Schrödinger equation gives the energies an electron can have. These are its energy levels. For each energy level, the Schrödinger equation also leads to a mathematical expression, called an atomic orbital, describing the probability of finding an electron at various locations around the nucleus. An **atomic orbital** is often thought of as a region of space in which there is a high probability of finding an electron.

The energy levels of electrons in the quantum mechanical model are labeled by principal quantum numbers (n). These are assigned the values $n = 1, 2, 3, 4$, and so forth. For each principal energy level, there may be several orbitals with different shapes and at different energy levels. These energy levels within a principal energy level constitute energy sublevels. **Each energy sublevel corresponds to an orbital of different shape describing where the electron is likely to be found.**

Different atomic orbitals are denoted by letters. As Figure 5.5 shows, *s* orbitals are spherical, and *p* orbitals are dumbbell-shaped. Because of the spherical shape of an *s* orbital, the probability of finding an electron at a given distance from the nucleus in an *s* orbital does not depend on direction. The three kinds of *p* orbitals have different orientations in space.

Checkpoint ***How do s and p orbitals differ?***

Interactive Textbook

Animation 5 Observe the characteristics of atomic orbitals.

with ChemASAP

Table 5.1

Summary of Principal Energy Levels, Sublevels, and Orbitals

Principal energy level	Number of sublevels	Type of sublevel
$n = 1$	1	1*s* (1 148rbital)
$n = 2$	2	2*s* (1 orbital), 2*p* (3 orbitals)
$n = 3$	3	3*s* (1 orbital), 3*p* (3 orbitals), 3*d* (5 orbitals)
$n = 4$	4	4*s* (1 orbital), 4*p* (3 orbitals), 4*d* (5 orbitals), 4*f* (7 orbitals)

Atomic Orbitals

Discuss L2

Present the orbital concept as an outgrowth of the quantum mechanical model and describe the shapes of the *s* and *p* orbitals. Stress that the quantum mechanical model predicts the shapes of the various orbitals. The fact that experiments affirm the model's predictions is validation that the model is correct. Remind students that an orbital is only a region of mathematical probability, not a concrete item that can be seen or felt. An orbital does not have the sharp boundaries that models and diagrams seem to indicate.

Use Visuals L1

Table 5.1 Explain that every atom has an infinite number of energy levels, but no element, when it is in its ground state, requires more than seven energy levels to describe the probability regions (orbitals) where its electrons can be found. Ask, **How many sublevels are in level 5?** (*5*) **in level 6?** (*6*) **in level 7?** (*7*). **Review the types of sublevels that are in level 5.** (*5s* (*1 orbital*), *5p* (*3 orbitals*), *5d* (*5 orbitals*), *5f* (*7 orbitals*) and *5g* (*9 orbitals*)) Ask, **What is the mathematical relationship between type (*s, p, d, f, g* and so on) and number of orbitals?** (*1, 3, 5, 7, 9 and so on*). Ask, **How many orbitals relate to level 1?** (*1*) **level 2?** (*4*) **level 3?** (*9*) **level 4?** (*16*) **Ask students if they see a pattern emerging.** (*The number of orbitals is the square of the energy level, n^2.*) **How many orbitals would you expect to relate to level 7?** (*49*) Stress that orbitals are mathematical probability regions where electrons may be located. The equation for determining how many electrons can relate to each energy level is $2n^2$ where n represents the energy level.

Differentiated Instruction

English Learners L1

Have students expand Table 5.1 to include a column with the heading *Number of orbitals (n^2)* and another headed *Number of electrons ($2n^2$)*. Have students fill in the rows. Make sure students do not confuse an orbital with any predictable physical location. Concentrate primarily on *s* and *p* orbitals because they are key to understanding the behavior of representative elements.

Answers to...

Checkpoint *S* orbitals are spherical, and *p* orbitals are dumbbell-shaped.

3 ASSESS

Evaluate Understanding L2

Ask students to discuss the parts of the Bohr model that are valid and the parts that are no longer considered accurate. Ask students to explain what a quantum of energy is and how an electron in an atom moves from one energy level to another.

Reteach L1

Review the following information. Energy levels are designated with the letter n. Each energy level has as many sublevels as the number of the energy level. When $n = 2$, there are 2 sublevels. Each energy level relates to n^2 orbitals. When $n = 3$, there are 3 sublevels and 9 orbitals ($1s + 3p + 5d = 9$). Each orbital contains up to 2 electrons. The maximum number of electrons at each energy level is $2n^2$. Thus, energy level 3 has up to 18 electrons.

Connecting Concepts

The Dalton model was a hard sphere; the Thomson model was a sphere of positive charge with embedded negative charges; the Rutherford model was a tiny nucleus containing positive charge with circulating electrons. The quantum mechanical model is of a tiny nucleus with electrons dispersed in a cloud surrounding it. The electron positions are dictated by probability.

Interactive Textbook

If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 5.1

with ChemASAP

Answers to...

Figure 5.6 Both lie in the xy plane. The lobes of the d_{xy} orbital lie between the x and y axes. Those of the $d_{x^2-y^2}$ orbital lie along the x and y axes.

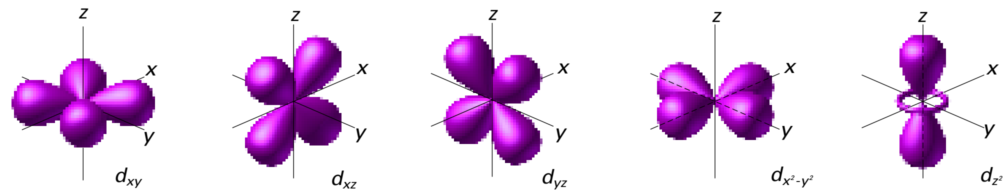


Figure 5.6 The d orbitals are illustrated here. Four of the five d orbitals have the same shape but different orientations in space. **Interpreting Diagrams** ***How are the orientations of the d_{xy} and $d_{x^2-y^2}$ orbitals similar? How are they different?***

Figure 5.6 shows the shapes of d orbitals. Four of the five kinds of d orbitals have clover leaf shapes. The shapes of f orbitals are more complicated than for d orbitals.

The numbers and kinds of atomic orbitals depend on the energy sublevel. The lowest principal energy level ($n = 1$) has only one sublevel, called $1s$.

The second principal energy level ($n = 2$) has two sublevels, $2s$ and $2p$. The $2p$ sublevel is of higher energy than the $2s$ and consists of three p orbitals of equal energy. The long axis of each dumbbell-shaped p orbital is perpendicular to the other two. It is convenient to label these orbitals $2p_x$, $2p_y$, and $2p_z$. Thus the second principal energy level has four orbitals: $2s$, $2p_x$, $2p_y$, and $2p_z$.

The third principal energy level ($n = 3$) has three sublevels. These are called $3s$, $3p$, and $3d$. As shown in Figure 5.6, the $3d$ sublevel consists of five d orbitals of equal energy. Thus the third principal energy level has nine orbitals (one $3s$, three $3p$, and five $3d$ orbitals).

The fourth principal energy level ($n = 4$) has four sublevels, called $4s$, $4p$, $4d$, and $4f$. The $4f$ sublevel consists of seven f orbitals of equal energy. The fourth principal energy level, then, has 16 orbitals (one $4s$, three $4p$, five $4d$, and seven $4f$ orbitals).

As mentioned, the principal quantum number always equals the number of sublevels within that principal energy level. The maximum number of electrons that can occupy a principal energy level is given by the formula $2n^2$, where n is the principal quantum number. The number of electrons allowed in each of the first four energy levels is shown in Table 5.2.

Table 5.2 Maximum Numbers of Electrons

Energy level n	Maximum number of electrons
1	2
2	8
3	18
4	32

5.1 Section Assessment

1. **Key Concept** Why did Rutherford's atomic model need to be replaced?
2. **Key Concept** What was the basic new proposal in the Bohr model of the atom?
3. **Key Concept** What does the quantum mechanical model determine about electrons in atoms?
4. **Key Concept** How do two sublevels of the same principal energy level differ from each other?
5. How can electrons in an atom move from one energy level to another?
6. The energies of electrons are said to be quantized. Explain what this means.
7. How many orbitals are in the following sublevels?
   a. $3p$ sublevel **b.** $2s$ sublevel **c.** $4p$ sublevel
   d. $3d$ sublevel **e.** $4f$ sublevel

Connecting Concepts

Reread the materials on the quantum mechanical model of the atom. Describe how the quantum mechanical model differs from Dalton's model, from Thomson's model, and from Rutherford's model.

Interactive Textbook

Assessment 5.1 Test yourself on the concepts in Section 5.1.

with ChemASAP

Section 5.1 Assessment

1. It could not explain why metals or compounds of metals give off characteristic colors when heated nor the chemical properties of the elements.
2. An electron is found only in specific circular paths or orbits around the nucleus.
3. It determines the allowed energy levels an electron can have and the likelihood of finding an electron in various locations around the nucleus.
4. The sublevels have different shapes.
5. by losing or gaining just the right amount of energy – a quantum
6. In an atom, the electrons can have certain fixed energy levels. To move from one energy level to another requires the emission or absorption of an exact amount of energy, or quantum. Thus the energy of the electron is said to be quantized.
7. **a.** 3 **b.** 1 **c.** 3 **d.** 5 **e.** 7