



CHAPTER 7

Electron Configurations and the Periodic Table

Objectives

You will be able to do the following.

1. Describe the two ways that scientists deal with the complexity and uncertainty associated with the modern understanding of the atom.
2. Describe how electrons are like radiant energy (light).
3. Describe the general steps associated with setting up and solving the wave equations for guitar strings and electrons.
4. Describe the information derived about guitar strings and electrons when their wave equations are solved.
5. Describe the $1s$ orbital in a hydrogen atom in terms of negative charge and in terms of the electron as a particle.
6. Explain why electrons in atoms are often described in terms of electron clouds.
7. Describe how electrons are like vibrating guitar strings.
8. Describe the information reflected by each of the four quantum numbers (n , l , m_l , and m_s) found in the electron wave equation.
9. Use the possible combinations of the quantum numbers n and l to show which sublevels exist for the first 4 principal energy levels for the one electron of a hydrogen atom.
10. Use the possible combinations of the quantum numbers n , l , and m_l , to show why s sublevels have one orbital, p sublevels have three orbitals, d sublevels have five orbitals, and f sublevels have seven orbitals for the one electron of a hydrogen atom.
11. Describe a $2s$ orbital for a hydrogen atom.
12. Explain why an electron in a hydrogen atom has lower potential energy in the $1s$ orbital than in the $2s$ orbital.
13. Describe the three $2p$ orbitals for a hydrogen atom.
14. Write or identify descriptions or drawings of the $3s$, $3p$, and $3d$ orbitals.
15. Write or identify the possible sublevels on the first seven principal energy levels.
16. Explain why hydrogen atoms heated to high temperature emit photons of only certain specific energies, wavelengths, and colors (for the visible portion of the spectrum).
17. Describe the effect on the intensity of light when light waves are in-phase and when they are out-of-phase.
18. Explain why when light of a specific wavelength is directed at a wall with two slits, a diffraction pattern, which consists of a series of lights and darks radiating out from the slits, is found on a wall opposite the slits.
19. Describe the difference between any two electrons in the same atomic orbital.

20. Describe the Pauli Exclusion Principle, and use it to explain (1) why an orbital can contain a maximum of two electrons, (2) why s sublevels can have a maximum of 2 electrons, (3) why p sublevels can have a maximum of 6 electrons, (4) why d sublevels can have a maximum of 10 electrons, and (5) why f sublevels can have a maximum of 14 electrons.
21. Explain why the $1s$ sublevel is filled before the $2s$ sublevel.
22. Explain why the $2s$ sublevel is filled before the $2p$ sublevel.
23. Explain why the $4s$ sublevel is filled before the $3d$ sublevel.
24. Draw a rough picture showing how the orbitals overlap for the electrons in a neon atom.
25. Write complete electron configurations, complete orbital diagrams, and abbreviated electron configurations for all of the elements on the periodic table that follow the normal order of filling of the sublevels.
26. Write complete electron configurations, orbital diagrams, and abbreviated electron configurations for atoms of copper, silver, gold, palladium, chromium, and molybdenum.
27. Identify atoms as either paramagnetic or diamagnetic.
28. Write complete electron configurations, orbital diagrams, and abbreviated electron configurations for monatomic ions.
29. Explain why transition metal atoms lose their $(n+1)s$ electrons before their nd electrons.
30. Explain why atoms of the following elements form the charges indicated: group 1 (+1), group 2 (+2), group 3 (+3), Cu^+ , Ag^+ , Au^+ , Zn^{2+} , Cd^{2+} , Hg^{2+} , Ga^+ , Ga^{3+} , In^+ , In^{3+} , Tl^+ , Tl^{3+} , Sn^{2+} , Pb^{2+} , Bi^{3+} , N^{3-} , P^{3-} , O^{2-} , S^{2-} , Se^{2-} , F^- , Cl^- , Br^- , and I^- .
31. Write or identify the trends for atomic size, first ionization energy, first electron affinity, and electronegativity within columns on the periodic table and within periods for the Representative elements, and within periods for the Transition Metals, and explain why these trends exist. Your explanation should include mention of the principle energy level for the highest energy electrons, nuclear charge, shielding of outer electrons by inner electrons, and effective charge.
32. Determine the effective charge for all the elements except the inner transition metals.
33. Given symbols for two elements and a periodic table, identify which would be expected to have atoms with the larger atomic size, higher ionization energy, more favorable electron affinity, and higher electronegativity.
34. Describe what it means for an ionic bond to have covalent character.
35. Given two pairs of atoms (one a metal and one a nonmetal in each case), predict which pair would form bonds with the most covalent character.
36. Given any atom or ion, write formulas for atoms or ions that form an isoelectronic series with it.
37. List the relative sizes of ions and atoms in an isoelectronic series and write an explanation for why they have this relative size.
38. Convert between the definition and the term for the following words or phrases.

Skip Sections 7.2, 7.3, 7.4, and 7.13.

Chapter 7 Glossary

Waveform A representation of the shape of a wave.

Standing (or stationary) wave A single frequency mode of vibration of a body or physical system in which the amplitude varies from place to place, is constantly zero at fixed points, and has maxima at other specific points.

Nodes The locations in a waveform where the intensity of the wave is always zero.

Orbitals The allowed waveforms for the electron in an atom. This term can also be defined as a volume that contains a high percentage of the electron charge or as a volume within which an electron has a high probability of being found.

Principal energy level or shell A collection of orbitals that have the same potential energy for a hydrogen atom, except for the first (lowest) principal energy level, which contains only one orbital ($1s$). For example, the $2s$ and $2p$ orbitals are in the second principal energy level.

Sublevel or Subshell A given type (or shape) of orbital available at a given principal energy level. For example, the second principal energy level contains a $2s$ sublevel (with one spherical orbital) and a $2p$ sublevel (with three dumbbell-shaped orbitals).

Ground state The condition of an atom whose electrons are in the orbitals that give it the lowest possible potential energy.

Excited state The condition of an atom that has at least one of its electrons in orbitals that do not represent the lowest possible potential energy.

Orbital diagram A drawing that uses lines or squares to show the distribution of electrons in orbitals and arrows to show the relative spin of each electron.

Electron configuration A description of the complete distribution of an element's electrons in atomic orbitals. Although a configuration can be described either with an orbital diagram or with its shorthand notation, this text will follow the common convention of referring to the shorthand notation that describes the distribution of electrons in sublevels without reference to the spin of the electrons as an electron configuration.

s block The portion of the periodic table for which the last electrons added to each element's electron configuration are added to an s orbital.

p block The portion of the periodic table for which the last electrons added to each element's electron configuration are added to an p orbital.

d block The portion of the periodic table for which the last electrons added to each element's electron configuration are added to an d orbital.

f block The portion of the periodic table for which the last electrons added to each element's electron configuration are added to an f orbital.

Paramagnetic Having a net magnetic field due to having at least one unpaired electron.

Diamagnetic Having no permanent net magnetic field due to having all electron paired.

van der Waals (or nonbonded radius) The radius of the sphere occupied by an atom with no covalent, ionic, or metallic bonds, e.g. one-half the distance between the nuclei of adjacent atoms for a noble gas in the solid form.

Ionic radius The radius of the sphere that an ion occupies.

Covalent radius The radius of the sphere that an ion occupies when it is covalently bonded to another atom. For some nonmetal atoms, it is one-half the distance between nuclei in their diatomic molecules.

Ionization energy The energy necessary to remove one mole of electrons from one mole of isolated and gaseous atoms or ions.

First ionization energy The energy necessary to remove one mole of electrons from one mole of isolated and gaseous uncharged atoms to form one mole of isolated and gaseous +1 ions.

Second ionization energy The energy necessary to remove one mole of electrons from one mole of isolated and gaseous +1 ions to form one mole of isolated and gaseous +2 ions.

Third ionization energy The energy necessary to remove one mole of electrons from one mole of isolated and gaseous +2 ions to form one mole of isolated and gaseous +3 ions.

Electron affinity The energy associated with adding one mole of electrons to one mole of isolated and gaseous atoms or ions.

First electron affinity The energy associated with adding one mole of electrons to one mole of isolated and gaseous uncharged atoms to form one mole of isolated and gaseous -1 ions.

Second electron affinity The energy associated with adding one mole of electrons to one mole of isolated and gaseous -1 ions to form one mole of isolated and gaseous -2 ions.

Third electron affinity The energy associated with adding one mole of electrons to one mole of isolated and gaseous -2 ions to form one mole of isolated and gaseous -3 ions.

Effective charge An approximate value for the charge felt by an atom's outer electron charge. It is an atom's nuclear charge minus the number of electron energy levels lower than the energy level for the highest energy electrons.

Isoelectronic series A collection of ions and an uncharged atom that have the same number of electrons.

One way to get a sense of the modern view of the electron is to compare it to light, or radiant energy. The following statements can be made about light.

“Although we know a great deal about *light*, we still have trouble describing what *it is*. *Light* seems to have a dual nature, with both particle and wave characteristics. It is difficult to describe these two aspects of *light* at the same time, so sometimes we focus on its particle nature and sometimes on its wave character, depending on which is more suitable in a given context.”

“In the particle view, *light is a stream of tiny, massless packets of energy called photons.*”

“In the wave view, as *light moves away from its source*, it has an effect on the space around it that can be described as a wave *consisting of an oscillating electric field perpendicular to an oscillating magnetic field.*”

“Because *light* seems to have both wave and particle characteristics, some experts have suggested that it is probably neither a wave nor a stream of particles. Perhaps the simplest model that includes both aspects of *light* says that as the

photons travel, they somehow affect the space around them in such a way as to create *the electric and magnetic fields.*”

With slight changes, these same statements can be made about electrons.

“Although we know a great deal about *electrons*, we still have trouble describing what *they are*. Each *electron* seems to have a dual nature, with both particle and wave characteristics. It is difficult to describe these two aspects of *an electron* at the same time, so sometimes we focus on its particle nature and sometimes on its wave character, depending on which is more suitable in a given context.”

“In the particle view, *electrons are tiny, negatively charged particles with a mass of about 9.1096×10^{-28} grams or 0.000549 atomic mass units.*”

“In the wave view, *an electron* has an effect on the space around it that can be described as a wave of *varying negative charge intensity.*”

“Because *an electron* seems to have both wave and particle characteristics, some experts have suggested that it is probably neither a wave nor a stream of particles. Perhaps the simplest model that includes both aspects of *electrons* says that as they travel, they somehow affect the space around them in such a way as to create a *three-dimensional wave of negative charge.*”

If you understand how guitar strings can be described in terms of wave mathematics, you will better understand the similar but much more complex mathematics of electrons. The following steps are used to develop descriptions of the possible waveforms for a vibrating guitar string.

STEP 1 Set up the general form of the wave equation that describes the vibrating string.

The following is the wave equation for a guitar string of length “a”. The equation describes the amplitude of the waveform in terms of the position x along the string. The amplitude at any point along the string is equal to the distance between the rest position of the string and the top (or bottom) of the waveform at that position.

$$A_x = A_0 \sin \frac{n\pi x}{a}$$

A_x = the amplitude at position x

A_0 = the maximum amplitude at any point on the string

n = 1, 2, 3,

x = the position along the string

a = the total length of the string

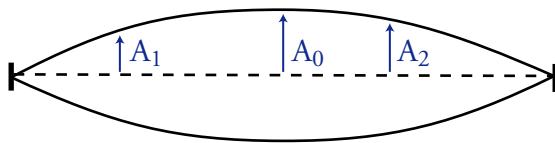


Figure 7.1

A_1 and A_2 represent amplitudes at $x = 1$ and $x = 2$.
 A_0 represents the maximum amplitude.

STEP 2 Determine the forms of the general equation that fit the boundary conditions.

Any possible vibration of the guitar string must have the ends of the string stationary. The equation above only leads to waveforms that have no movement at the ends of the string when n in the equation is an integer value. For the guitar string, different n values lead to different possible equations. (For n equals 1, we get one allowed equation that meets the boundary conditions, for n equals 2 we get another equation, etc.)

$$A_x = A_0 \sin \frac{\pi x}{a} \quad A_x = A_0 \sin \frac{2\pi x}{a} \quad A_x = A_0 \sin \frac{3\pi x}{a}$$

STEP 3 Each possible equation is solved over and over again for the amplitudes at many different positions.

We could first use the following equation to calculate amplitude values for one waveform. For example, let's assume that we have a guitar string that is one meter long and that we pluck it hard enough to cause it to vibrate with a maximum amplitude of 5 cm (or 0.05 m). If we plug in 0.05 m for A_0 and 1 m for "a", our first allowed equation simplifies to the second form shown below.

$$A_x = A_0 \sin \frac{\pi x}{a} \quad A_x = 0.05 \sin \pi x$$

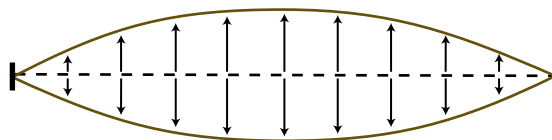
Now we could calculate the amplitude at many different positions along the string. For example, we could calculate the amplitude every 0.10 m. We could calculate A_x at $x = 0.10$ m, then repeat the calculation at $x = 0.20$ m, etc.

STEP 4 We plot the values determined in Step 3 and get an image of the possible wave forms.

We could plot the values for the first of our allowed equations to see the allowed waveform it predicts. See Figure 7.2.

Figure 7.2

One Allowed Waveform for a Guitar String
The arrows represent the amplitudes that are calculated at various positions on the string.



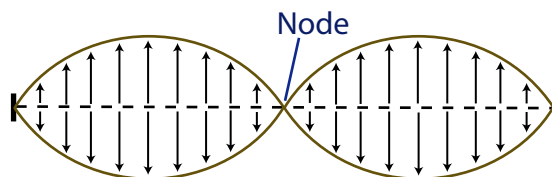
Step 5 Steps 3 and 4 can be repeated for other equations that meet the boundary conditions.

Calculations of the amplitudes at various positions along the string for the second possible equation for our guitar string (with $n = 2$) leads to the waveform in Figure 7.3.

$$A_x = A_0 \sin \frac{2\pi x}{a}$$

Figure 7.3

Second Allowed Waveform for a Guitar String
The arrows represent the amplitudes that are calculated at various positions on the string.



Similar calculations for the third possible equation lead to a different waveform (Figure 7.4).

$$A_x = A_0 \sin \frac{3\pi x}{a}$$

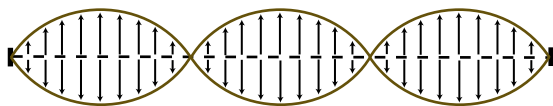


Figure 7.4

Third Allowed Waveform for a Guitar String
The arrows represent the amplitudes that are calculated at various positions on the string.

The following statements represent the core of the modern description of the wave character of the electron:

- Just as the intensity of the movement of a guitar string can vary, so can the intensity of the negative charge of the electron vary at different positions outside the nucleus.
- The variation in the intensity of the electron charge can be described in terms of a three-dimensional standing wave *like* the standing wave of the guitar string.
- As in the case of the guitar string, only certain waveforms are possible for the electron in an atom.
- We can focus our attention on the waveform of varying intensity of movement of a guitar string (or varying intensity of electron charge) without having to think about the actual physical nature of the guitar string (or electron).

The following steps can be taken to determine the possible waveforms for the one electron in a hydrogen atom.

Step 1 Set up the general form of the wave equation that describes the electron.

The amplitude for an allowed waveform for a guitar string varies at different positions along the string. We can calculate the amplitude, A_x , using the guitar string's wave equation by plugging in various values that represent the position "x" along the string. We say that the amplitude of a guitar string is a function of the position x along the string. The shorthand for this is below.

$$A_x = f(x)$$

The intensity of the wave character of an electron varies at different positions outside the nucleus. Positions outside the nucleus can be described in terms of their x, y, and z coordinates in a three-dimensional coordinate system with the nucleus at the center. The wave equation for the electron has the following form.

$$\Psi_{x,y,z} = f(x,y,z)$$

An equation that describes the wave character of an electron is called a *wave function*. The values for the wave function, Ψ , are *like* the values for the amplitude of a vibrating guitar string. They both lead to descriptions of the intensity of the wave character at various positions. For the guitar string, this wave character is related to the movement of the string at various positions along the string. A large amplitude at a position on a guitar string suggests a lot of movement.

If we focus our attention only on the electron charge and not the particle nature of the electron, the values for the electron wave equation relate to the intensity (or concentration) of the negative charge of the electron at various positions outside the nucleus. A large value for Ψ at a particular position outside the nucleus suggests a high concentration of negative charge at that position.

Step 2 Determine the forms of the general equation that fit the boundary conditions.

The boundary conditions for the electron mathematics are more complex than those for the guitar string, but when they are applied, we find more than one possible wave function. These lead to the prediction of different possible waveforms for the electron that are like the different possible waveforms for the vibrating guitar string. Although there are an infinite number of possible wave functions for an electron in a hydrogen atom, we are only concerned with eighteen of them. Each of these wave functions is described with a symbol such as $1s$, $2s$, or $2p$.

$$\Psi_{1s} = f_{1s}(x,y,z) \quad \Psi_{2s} = f_{2s}(x,y,z) \quad \Psi_{2p} = f_{2p}(x,y,z)$$

STEP 3 Each allowed equation is solved to get the values for the wave function for many different positions.

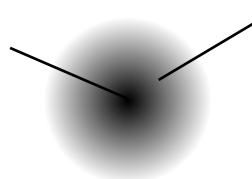
For example, the first of the wave functions above, Ψ_{1s} , can be solved over and over for various positions outside the nucleus. The values at each position reflect the intensity (or concentration) of negative charge at that point. The best mathematical description of the intensity of the negative charge actually comes from the square of the values for the wave function, Ψ^2 .

STEP 4 When we plot the values for one of the possible wave functions on a three-dimensional coordinate system, we get an image of one of the possible waveforms.

Figure 7.5

Waveform of the $1s$ Electron

Nucleus, about 0.000001 the diameter of the atom



The negative charge is most intense at the nucleus and decreases in intensity with distance outward.

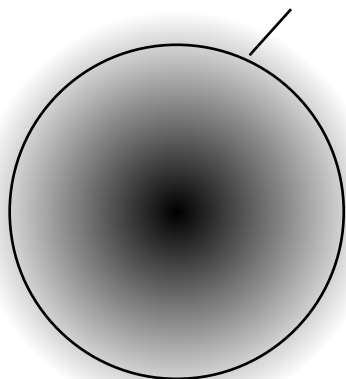
The mathematics predicts that the intensity of the negative charge created by a $1s$ electron in a hydrogen atom approaches zero with increasing distance from the nucleus, but it never gets there. Although it might be amusing to consider the suggestion that some of the negative charge created by an electron in a hydrogen atom is felt an infinite distance from the atom's nucleus, it is more useful to find a volume that contains most of the electron charge and focus our attention on this volume, forgetting about the small negative charge felt outside this volume. For example, we can draw a spherical surface around 90% of the charge of the $1s$ electron. If we wanted to include more of the electron charge, we could draw a larger surface that contains 99% (or 99.9%) of the electron charge

The allowed waveforms for the electron can be called **orbitals**. The term orbital is defined in several ways. It is sometimes defined as the wave function itself, Ψ . An orbital can also be defined as the volume which contains a high percentage of the

electron charge.

Most of the pictures that you might see of orbitals represent the surfaces that surround a high percentage of the negative charge for the electron that has that waveform. The 1s orbital can either be represented by a fuzzy picture or a drawing of the surface that contains most of its negative charge.

Almost all of the electron's charge lies within a spherical shell with the diameter of this circle.



Sphere enclosing almost all of the electron's negative charge

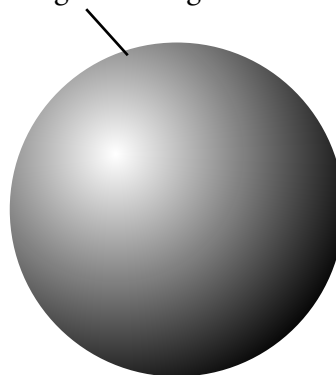


Figure 7.6
1s Orbital

Is the sphere in the Figure 7.6 above the 1s electron? This is similar to asking if light is its oscillating electric and magnetic fields or asking if the guitar string is the blur that you see when you squint at the vibrating string. When we describe the wave character of light, we usually do not consider the photons. The standing wave that represents the motion of the guitar string can be described without reference to the material of the guitar string. The situation is very similar for the electron. We are able to describe the variation of negative charge created by the electron without thinking too much about what the electron is and what it is doing.

Other Important Waveforms

Just like the guitar string can have different waveforms, the one electron in a hydrogen atom can also have different waveforms, or orbitals. The shapes and sizes for these orbitals are predicted by the mathematics associated with the wave character of the hydrogen electron. Figure 7.7 shows some of them.

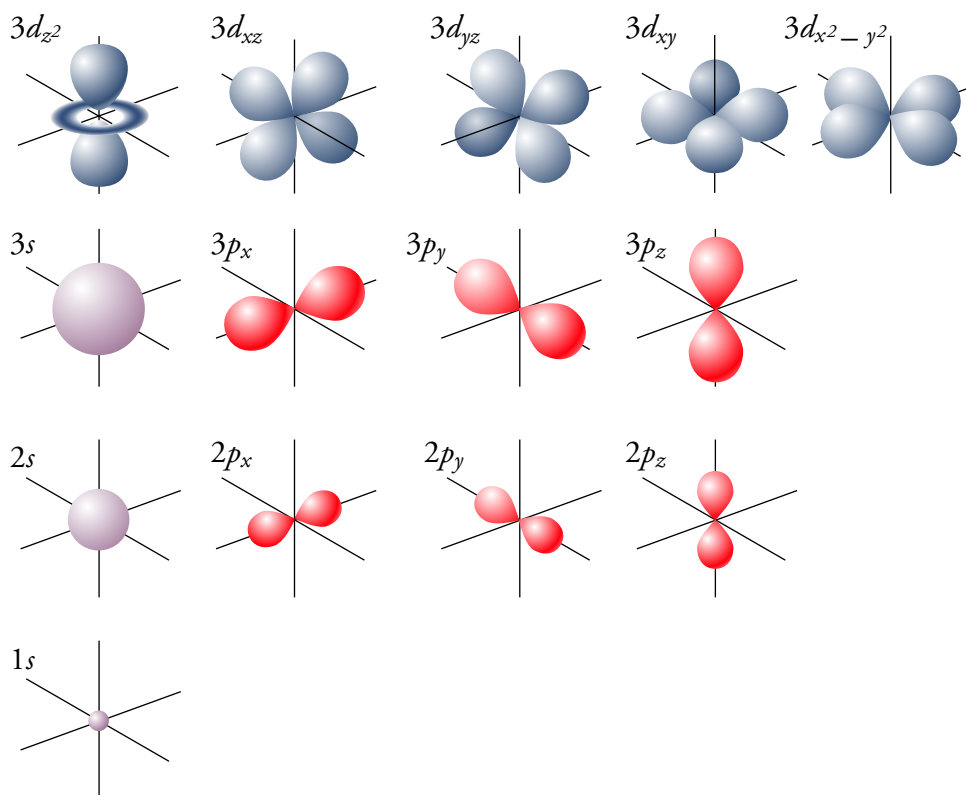
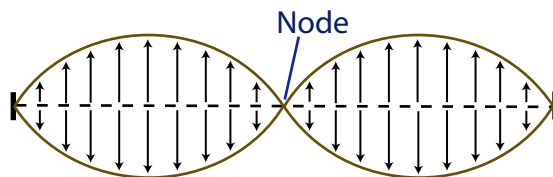


Figure 7.7
Some Possible
Waveforms or Orbitals
for an Electron in a
Hydrogen Atom

Before considering the second possible orbital for the electron of a hydrogen atom, let's look at another of the possible ways a guitar string can vibrate. The guitar-string waveform below has a node in the center where there is no movement of the string.



The electron-wave calculations predict that an electron in a hydrogen atom can have a waveform called the $2s$ orbital that is analogous to the guitar-string waveform above. The $2s$ orbital for an electron in a hydrogen atom is spherical like the $1s$ orbital, but it is a larger sphere. All spherical electron waveforms are called s orbitals. For an electron in the $2s$ orbital, the charge is most intense at the nucleus. With increasing distance from the nucleus, the charge diminishes in intensity until it reaches a minimum at a certain distance from the nucleus; it then increases again to a maximum, and finally it diminishes again. The region within the $2s$ orbital where the charge intensity decreases to zero is called a **node**. Figure 7.8 shows cutaway, quarter-section views of the $1s$ and $2s$ orbitals.

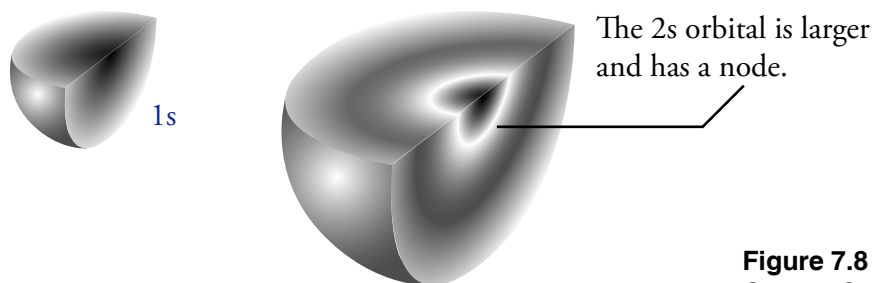
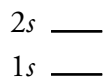


Figure 7.8
Quarter Sections of the 1s and 2s
Orbitals

The average distance between the positive charge of the nucleus and the negative charge of a $2s$ electron cloud is greater than the average distance between the nucleus and the charge of a $1s$ electron cloud. Because the strength of the attraction between positive and negative charges decreases with increasing distance between the charges, an electron is more strongly attracted to the nucleus and therefore is more stable when it has the smaller $1s$ waveform than when it has the larger $2s$ waveform. Increased stability is associated with decreased potential energy, so a $1s$ electron has lower potential energy than a $2s$ electron. We describe this energy difference by saying the $1s$ electron is in the first principal energy level, and the $2s$ electron is in the second principal energy level. All of the orbitals that have the same potential energy for a hydrogen atom are said to be in the same **principal energy level**. The principal energy levels are often called **shells**. The 1 in $1s$ and the 2 in $2s$ show the principal energy levels, or shells, for these orbitals.

Chemists sometimes draw orbital diagrams, such as the following, with lines to represent the orbitals in an atom and arrows (which we will be adding later) to represent electrons:



The line representing the $2s$ orbital is higher on the page to indicate its higher potential energy.

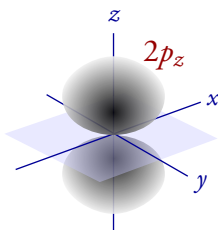
Because electrons seek the lowest energy level possible, we expect the electron in a hydrogen atom to have the $1s$ waveform, or electron cloud. We say that the electron is *in* the $1s$ orbital. But the electron in a hydrogen atom does not need to stay in the $1s$ orbital at all times. Just as an input of energy (a little arm work on our part) can lift a book resting on a table and raise it to a position that has greater potential energy, so can the waveform of an electron in a hydrogen atom be changed from the $1s$ shape to the $2s$ shape by the addition of energy to the atom. We say that the electron can be *excited* from the $1s$ orbital to the $2s$ orbital. Hydrogen atoms with their electron in the $1s$ orbital are said to be in their **ground state**. A hydrogen atom with its electron in the $2s$ orbital is in an **excited state**.

If you lift a book from a table to above the table and then release it, it falls back down to its lower-energy position on the table. The same is true for the electron. After the electron is excited from the $1s$ orbital to the $2s$ orbital, it spontaneously returns to its lower-energy $1s$ form.

An electron in a hydrogen atom can be excited to orbitals other than the $2s$. For

example, an electron in a hydrogen atom can be excited from the $1s$ to a $2p$ orbital. There are actually three possible $2p$ orbitals. They are identical in shape and size, but each lies at a 90° angle to the other two. Because they can be viewed as lying on the x , y and z axes of a three-dimensional coordinate system, they are often called the $2p_x$, $2p_y$, and $2p_z$ orbitals. An electron with a $2p$ waveform has its negative charge distributed in two lobes on opposite sides of the nucleus. Figure 7.9 shows the shape of a $2p_z$ orbital.

Figure 7.9
Realistic View of a
 $2p_z$ Orbital



In order to more easily show how the $2p$ orbitals fit together, they are often drawn in a more elongated and stylized form (Figures 7.10 and 7.11).

Figure 7.10
Stylized View of a
 $2p_z$ Orbital

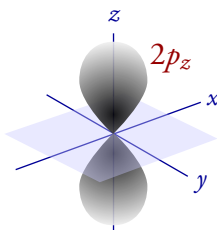
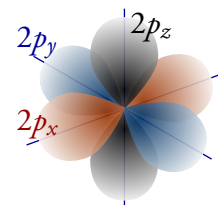
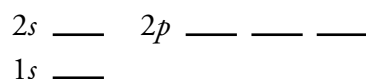


Figure 7.11
Stylized View of the
 $2p_x$, $2p_y$, and $2p_z$
Orbitals Combined



In a hydrogen atom, the average distance between the negative charge of a $2p$ electron cloud and the nucleus is the same as for a $2s$ electron cloud. Therefore, an electron in a hydrogen atom has the same attraction to the nucleus and the same stability when it has a $2p$ form as when it has the $2s$ form. Therefore, a $2s$ electron has the same potential energy as a $2p$ electron. These orbitals are in the same principal energy level.

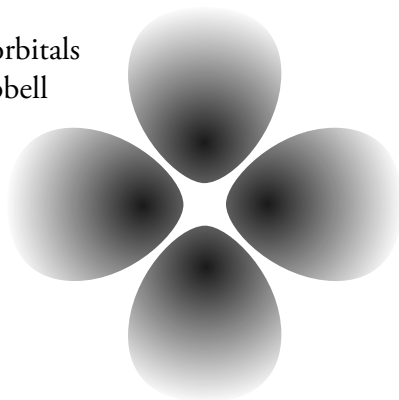
Because the shapes of the $2s$ and $2p$ electron clouds are different, we distinguish between them by saying that an electron with the $2s$ waveform is in the $2s$ sublevel, and an electron with any of the three $2p$ waveforms is in the $2p$ sublevel. Orbitals that have the same potential energy, the same size, and the same shape are in the same **sublevel**. The sublevels are sometimes called **subshells**. Thus there is one orbital in the $1s$ sublevel, one orbital in the $2s$ sublevel, and three orbitals in the $2p$ sublevel. The following orbital diagram shows these orbitals and sublevels.



The lines for the $2s$ and $2p$ orbitals are drawn side-by-side to show that they have the same potential energy for the one electron in a hydrogen atom. You will see in the next section that the $2s$ and $2p$ orbitals have different potential energies for atoms larger than hydrogen atoms.

In the third principal energy level, there are nine possible waveforms for an electron, in three different sublevels. The $3s$ sublevel has one orbital that has a spherical shape, like the $1s$ and the $2s$, but it has a larger average radius and two nodes. Its greater average distance from the positive charge of the nucleus makes it less stable and higher in energy than the $1s$ or $2s$ orbitals. The calculations predict that the third principal energy level has a $3p$ sublevel with three $3p$ orbitals. They have the same general shape as the $2p$ orbitals, but they are larger, which leads to less attraction to the nucleus, less stability, and higher potential energy than for a $2p$ orbital. The third principal energy level also has a $3d$ sublevel with five $3d$ orbitals. Four of these orbitals have four lobes whose shape is similar to the lobes of a $3p$ orbital. We will call these “double dumbbells.” An electron in a $3d$ orbital has its negative charge spread out in these four lobes. The fifth $3d$ orbital has a different shape, as shown in Figure 7.12.

Four of the five $3d$ orbitals have a double dumbbell shape like this one.



The fifth $3d$ orbital is shaped like a dumbbell and a donut.

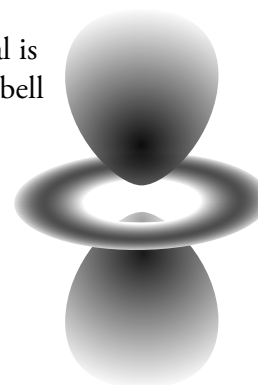


Figure 7.12
3d Orbitals

When an electron in a hydrogen atom is excited to the fourth principal energy level, it can be in any one of four sublevels: $4s$, $4p$, $4d$, or $4f$. There is one $4s$ orbital, with a spherical shape and a larger volume than the $3s$. There are three $4p$ orbitals, similar in shape to the $3p$ orbitals but with a larger volume than the $3p$ orbitals. There are five $4d$ orbitals, similar to but larger than the $3d$ orbitals. The $4f$ sublevel has seven possible orbitals.

Overall Organization of Principal Energy Levels, Sublevels, and Orbitals

Table 7.1 shows all the orbitals predicted for the first seven principal energy levels. Notice that the first principal energy level has one sublevel, the second has two sublevels, the third has three sublevels, and the fourth has four. If “ n ” is the number associated with the principal energy level, each principal energy level has “ n ” sublevels. Thus there are five sublevels on the fifth principal energy level: $5s$, $5p$, $5d$, $5f$, and $5g$. The $5s$, $5p$,

$5d$, and $5f$ orbitals have shapes similar to the $4s$, $4p$, $4d$, and $4f$ orbitals, but they are larger and have higher potential energy.

Each s sublevel has one orbital, each p sublevel has three orbitals, each d has five orbitals, and each f sublevel has seven orbitals. Thus there are one $5s$ orbital, three $5p$ orbitals, five $5d$ orbitals, and seven $5f$ orbitals. The trend of increasing the number of orbitals by two for each succeeding sublevel continues for $5g$ and beyond. There are nine $5g$ orbitals with shapes more complex than the shapes of the $4f$ orbitals.

Table 7.1 Possible Sublevels and Orbitals for the First Seven Principal Energy Levels
The sublevels in parentheses are not necessary for describing any of the known elements.

Principal Energy Level (Shell)	Sublevels (Subshells)	Number of Orbitals
1	$1s$	1
2	$2s$	1
	$2p$	3
3	$3s$	1
	$3p$	3
	$3d$	5
4	$4s$	1
	$4p$	3
	$4d$	5
	$4f$	7
5	$5s$	1
	$5p$	3
	$5d$	5
	$5f$	7
	$(5g)$	9
6	$6s$	1
	$6p$	3
	$6d$	5
	$(6f)$	7
	$(6g)$	9
	$(6h)$	11
7	$7s$	1
	$7p$	3
	$(7d)$	5
	$(7f)$	7
	$(7g)$	9
	$(7h)$	11
	$(7i)$	13

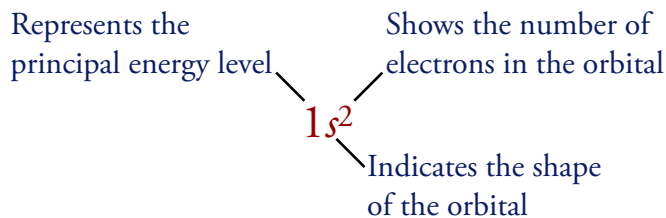


Figure 7.13
Electron Configuration

The procedure for describing the electron configurations and orbital diagrams for elements beyond helium is guided by the following three principles.

- The sublevels are filled in such a way as to yield the lowest overall potential energy for the atom.
- No two electrons in an atom can be the same in all ways. As we will see, this is one way to describe the Pauli exclusion principle.
- When electrons are filling orbitals of the same energy, they prefer to enter empty orbitals first, and all electrons in half-filled orbitals have the same spin. This is called Hund's Rule.

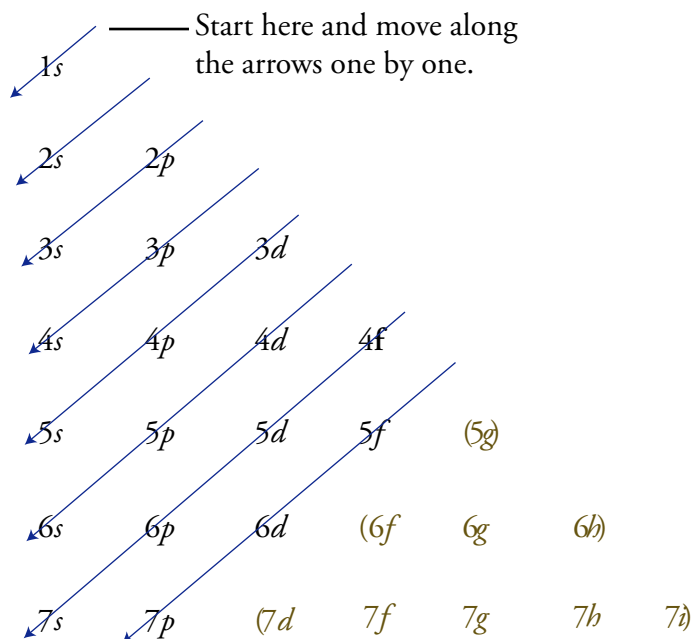


Figure 7.14
Aid for remembering the Order of Sublevel Filling

No two electrons in an atom can be the same in all ways. This is called the Pauli exclusion principle. There are four ways that electrons can be the same:

- Electrons can be in the same principal energy level.
- They can be in the same sublevel.
- They can be in the same orbital.
- They can have the same spin.

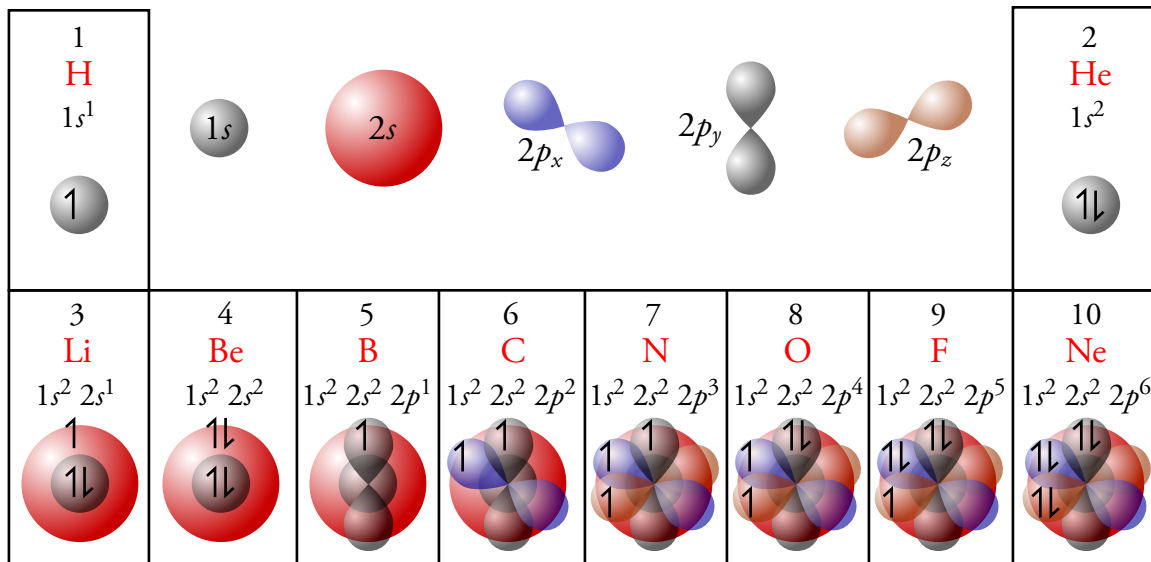


Figure 7.15
Electron Configurations and Orbital Models for the First Ten Elements

Table 7.2 Numbers of Orbitals and Electrons per Sublevel

Type of Sublevel	Number of Orbitals	Maximum Number of Electrons
<i>s</i>	1	2
<i>p</i>	3	6
<i>d</i>	5	10
<i>f</i>	7	14

Figure 7.16
Periodic Table with the Inner Transition Metals in their Natural Position

Sample Study Sheet 7.1: Complete Electron Configurations and Orbital Diagrams for Uncharged Atoms

TIP-OFF If you are asked to write a complete electron configuration or an orbital diagram, you can use the following guidelines.

GENERAL STEPS -

To write a complete electron configuration for an uncharged atom,

- Determine the number of electrons in the atom from its atomic number.
- Add electrons to the sublevels in the correct order of filling.

Add two electrons to each *s* sublevel, 6 to each *p* sublevel, 10 to each *d* sublevel, and 14 to each *f* sublevel.

- To check your complete electron configuration, look to see whether the location of the last electron added corresponds to the element's position on the periodic table.

To draw an orbital diagram for an uncharged atom,

- Write the complete electron configuration for the atom. (This step is not absolutely necessary, but it can help guide you to the correct orbital diagram.)
- Draw a line for each orbital of each sublevel mentioned in the complete electron configuration.

Draw one line for each *s* sublevel, three lines for each *p* sublevel, five lines for each *d* sublevel, and seven lines for each *f* sublevel.

As a guide to the order of filling, draw your lines so that the orbitals that fill first are lower on the page than the orbitals that fill later.

Label each sublevel.

- For orbitals containing two electrons, draw one arrow up and one arrow down to indicate the electrons' opposite spin.
- For unfilled sublevels, add electrons to empty orbitals whenever possible, giving them the same spin.

The arrows for the first three electrons to enter a *p* sublevel should each be placed pointing up in different orbitals. The fourth, fifth, and sixth are then placed, pointing down, in the same sequence, so as to fill these orbitals.

The first five electrons to enter a *d* sublevel should be drawn pointing up in different orbitals. The next five electrons are drawn as arrows pointing down and fill these orbitals (again, following the same sequence as the first five *d* electrons).

The first seven electrons to enter an *f* sublevel should be drawn as arrows pointing up in different orbitals. The next seven electrons are paired with the first seven (in the same order of filling) and are drawn as arrows pointing down.

EXERCISE 7.1 - Complete Electron Configuration and Orbital Diagram

Write the complete electron configuration and orbital diagram for antimony, Sb.

EXERCISE 7.2 - Abbreviated Electron Configurations

Write the abbreviated electron configurations for each of the following.

a. rubidium, Rb

b. nickel, Ni

c. bismuth, Bi

Some of the elements have electron configurations that differ slightly from what our general procedure would lead us to predict. You can read more about these at the following web address:

www.mpcfacy.net/mark_bishop/anomalies.htm

The following web address shows you how to predict charges on monatomic ions and how to write their electron configurations:

www.mpcfacy.net/mark_bishop/monatomic_ion_configurations.htm

Table 7.3 Unusual Electron Configurations

Element	Predicted Electron Configuration	Actual Electron Configuration
copper, Cu	[Ar] $3d^9 4s^2$	[Ar] $3d^{10} 4s^1$
silver, Ag	[Kr] $4d^9 5s^2$	[Kr] $4d^{10} 5s^1$
gold, Au	[Xe] $4f^{14} 5d^9 6s^2$	[Xe] $4f^{14} 5d^{10} 6s^1$
palladium, Pd	[Kr] $4d^8 5s^2$	[Kr] $4d^{10}$
chromium, Cr	[Ar] $3d^4 4s^2$	[Ar] $3d^5 4s^1$
molybdenum, Mo	[Kr] $4d^4 5s^2$	[Kr] $4d^5 5s^1$

1		2				13		14	15	16	17	18	
1A		2A				3A		4A	5A	6A	7A	8A	
$1s^2$ Li ⁺	$1s^2$ Be ²⁺			$1s^2$ H ⁻					$2s^2 2p^6$ N ³⁻	$2s^2 2p^6$ O ²⁻	$2s^2 2p^6$ F ⁻	$2s^2 2p^6$ Ne	
$2s^2 2p^6$ Na ⁺	$2s^2 2p^6$ Mg ²⁺					$2s^2 2p^6$ Al ³⁺			$3s^2 3p^6$ P ³⁻	$3s^2 3p^6$ S ²⁻	$3s^2 3p^6$ Cl ⁻	$3s^2 3p^6$ Ar	
$3s^2 3p^6$ K ⁺	$3s^2 3p^6$ Ca ²⁺					$3d^{10}$ Cu ⁺	$3d^{10}$ Zn ²⁺	$3d^{10} 4s^2$ Ga ⁺	$3d^{10}$ Ga ³⁺		$4s^2 4p^6$ Se ²⁻	$4s^2 4p^6$ Br ⁻	$4s^2 4p^6$ Kr
$4s^2 4p^6$ Rb ⁺	$4s^2 4p^6$ Sr ²⁺					$4d^{10}$ Ag ⁺	$4d^{10}$ Cd ²⁺	$4d^{10} 5s^2$ In ⁺	$4d^{10}$ In ³⁺	$4d^{10} 5s^2$ Sn ²⁺		$5s^2 5p^6$ I ⁻	$5s^2 5p^6$ Xe
$5s^2 5p^6$ Cs ⁺	$5s^2 5p^6$ Ba ²⁺					$4f^{14} 5d^{10}$ Au ⁺	$4f^{14} 5d^{10}$ Hg ²⁺	$4f^{14} 5d^{10} 6s^2$ Tl ⁺	$4f^{14} 5d^{10}$ Tl ³⁺	$4f^{14} 5d^{10} 6s^2$ Pb ²⁺	$4f^{14} 5d^{10} 6s^2$ Bi ³⁺		$6s^2 6p^6$ Rn
$6s^2 6p^6$ Fr ⁺	$6s^2 6p^6$ Ra ²⁺					$6s^2 6p^6$ Lu ³⁺							$6s^2 6p^6$ Lr ³⁺

Figure 7.19 Ions with Predictable Charges
Some of these elements also form other, less easily predicted charges.

TIP-OFF – If you are asked to predict the charge or charges on monatomic ions of an element, or if you are asked to write abbreviated electron configurations for monatomic ions, follow these steps.

GENERAL STEPS

To predict ionic charges, follow these guidelines.

- Nonmetallic elements form anions.

Hydrogen atoms gain one electron to form H⁻ with a stable $1s^2$ electron configuration.

The other nonmetallic elements gain one, two, or three electrons to achieve a stable $ns^2 np^6$ configuration. The group 17 nonmetals form -1 ions, the group 16 nonmetals form -2 ions, and the group 15 nonmetals form -3 ions.

- Metallic elements form cations. (You will only be able to predict some of the charges on these cations.)

Many elements lose one, two, or three electrons and achieve one or two of the stable electron configurations: $1s^2$, $ns^2 np^6$, nd^{10} , or $nd^{10}(n+1)s^2$.

The exact procedure for writing the abbreviated electron configuration for an ion depends on whether the ion is an anion; Al³⁺ or a cation in groups 1, 2, or 3; or any other cation.

Sample Study Sheet 7.3: Predicting Ionic Charges and Writing Abbreviated Electron Configurations for Monatomic Ions

For monatomic anions, follow these steps.

- Locate the symbol for the element on the periodic table.
- Move to the far right of the same row on the table to find the nearest larger noble gas.
- Write the abbreviated electron configuration for this noble gas.

This can be done most simply by putting the symbol for this noble gas in brackets, For example, [Ar] for argon.

It can also be done using the symbol for the noble gas element at the end of the previous row. For example, [Ne] $3s^2 3p^6$ for argon.

For Al^{3+} and monatomic cations from elements in groups 1, 2, or 3, follow these steps.

- Locate the symbol for the element on the periodic table.
- Move up one row and to the far right on the table to find the symbol for the noble gas to put in brackets.
- Write the abbreviated electron configuration for this noble gas.

For any other monatomic cations, follow these steps.

- Write the abbreviated electron configuration for the uncharged atom, listing the sublevels for the electrons outside of the noble gas configuration in the order of increasing principal energy level.
- Remove electrons from the electron configuration of the uncharged atom starting with the electrons listed on the far right. Remove one electron for +1 cations, two electrons for +2, and three electrons for +3.

EXERCISE 7.3 - Abbreviated Electron Configurations for Monatomic Ions

Write the abbreviated electron configurations for each of the following cations.

- a. cadmium ion, Cd^{2+}
- b. lead(II) ion, Pb^{2+}

EXERCISE 7.4 - Predicting and Explaining Monatomic Ion Charges

Predict the charge or charges for the ions that atoms of each of the following elements would form. Write electron configurations for each ion and use it to explain why the ion forms the charge or charges it does.

- a. selenium, Se
- b. yttrium, Y
- c. thallium, Tl

EXERCISE 7.5 - Effective Charge

What is the effective charge on the highest energy electrons of atoms of each of the following elements?

- a. Na b. Al c. Sc d. Ni

Effective charge as we calculate it is an approximation. It assumes that the electrons in lower energy levels shield outer electrons 100%. Due to the penetration of the inner core of electrons, the shielding is not 100%.

- A better assumption is 15% shielding by electrons on the same energy level.
- A better value for shielding from electrons in an energy level one below the outer level is 85% shielding.
- It is a good approximation to assume that electrons in energy levels two or more lower shield 100%.

EXERCISE 7.6 - Atomic size and Ionization Energy

Identify the element in each pair below that has atoms with larger atomic sizes and higher ionization energy. If they are about the same, write “neither”.

- a. S or Se
b. C or F
c. Ni or Pt
d. Mn or Co

EXERCISE 7.7 - Electron Affinity

Circle the symbol in each pair that represents the element that has the most favorable electron affinity.

- a. Cl or I b. C or F

EXERCISE 7.8 - Isoelectronic Series

Write the formulas for three cations and three anions that are isoelectronic with krypton and arrange them in the order of increasing ionic size.

