

# The Electron Configuration of Atoms

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# Chapter 1

## The Electron Configuration of Atoms

### 1.1 Electron Arrangement

#### Lesson Objectives

The student will:

- draw the orbital representation for selected atoms.
- write the electron configuration code for selected atoms.
- identify the principal, angular momentum, magnetic, and spin quantum numbers.
- identify the four quantum numbers for indicated electrons.
- identify the energy level, sub-energy level, orbital, and spin for an electron given the four quantum numbers for the electron

#### Vocabulary

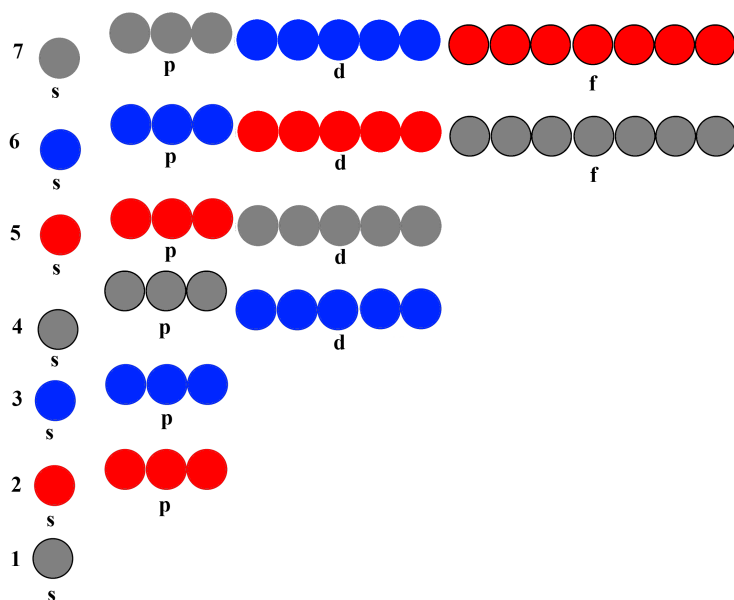
- angular momentum quantum number
- Aufbau principle
- electron configuration code
- Hund's rule
- magnetic quantum number
- orbital representation
- spin quantum number

#### Introduction

When dealing with the chemical behavior of atoms, it is not feasible to consider electrons as standing waves, nor is it feasible to indicate the positions of electrons in the electron cloud by drawing probability patterns. Fortunately, chemists have developed simpler and faster methods of describing the ways that electrons are arranged within atoms. It is important to be familiar with all of these methods, because the most useful way to represent the electron distribution often depends on the particular situation.

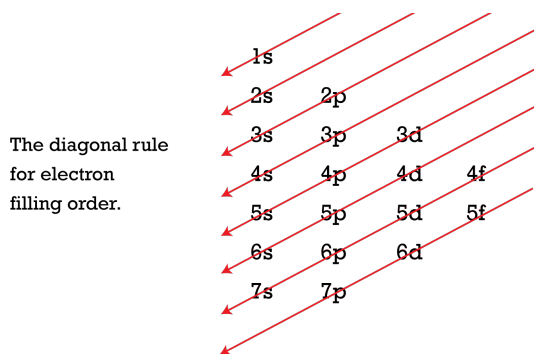
# The Orbital Representation

The **orbital representation** uses circles or lines to represent each orbital. In the figure below, all the energy levels, sub-levels, and orbitals are represented through  $7p$ . The distance from the bottom of the chart indicates the energy of each energy level and sub-level. The closer the energy level is to the bottom of the chart, the lower its energy. At the bottom of the chart, you will find the first energy level,  $n = 1$ . The chart shows only one circle in the first energy level. This is because the first energy level contains only one  $s$  sub-level, and  $s$  sub-levels have only one orbital.



Proceeding up the chart (higher energy), we see the second energy level,  $n = 2$ . The second energy level has both an  $s$  sub-level and a  $p$  sub-level. The  $s$  sub-level in the second energy level, like all  $s$  sub-levels, contains only one orbital and is represented by a single circle. The second energy level also has a  $p$  sub-level, which consists of three orbitals. The three orbitals in a  $p$  sub-level all have the same energy. Therefore, the  $p$  sub-level is represented by three circles equidistant from the bottom of the chart. This process is continued through the first seven energy levels and each of the sub-levels.

Careful observation of the orbital representation chart will show that the next higher energy sub-level is not always what you might expect. If you look closely at the relationship between the  $3d$  and  $4s$  sub-levels, you will notice that the  $4s$  sub-level is slightly lower in energy than the  $3d$  sub-level. Therefore, when the energy sub-levels are being filled with electrons, the  $4s$  orbital is filled before the  $3d$  orbitals. As you go up the chart, there are more of these variations. The complete filling order is given in the chart below.



Filling order:  $1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p$

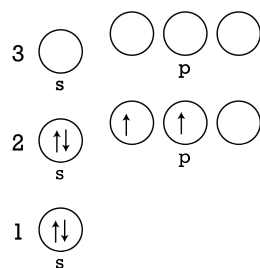
To determine the filling order, follow the top arrow from the base to the arrowhead. When you reach the arrowhead, move to the base of the next arrow and follow it to the arrowhead.

This process of building up the atoms by adding one proton to the nucleus and one electron to the electron cloud is known as the Aufbau principle. The **Aufbau principle** states that as electrons are added to “build up” the elements, each electron is placed in the lowest energy orbital available. (*Aufbau* is German for “building up”.)

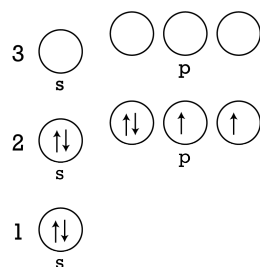
### Rules for Determining Electron Configuration:

1. Following the Aufbau principle, each added electron enters the lowest energy orbital available.
2. No more than two electrons can be placed in any orbital.
3. Before a second electron can be placed in any orbital, all the orbitals of that sub-level must contain at least one electron. This is known as **Hund’s Rule**.

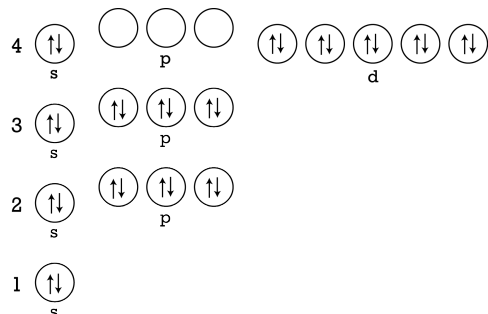
Here is the orbital representation for the electron configuration of carbon:



A carbon atom has six protons and six electrons. When placing the electrons in the appropriate orbitals, the first two electrons go into the 1s orbital, the second two go into the 2s orbital, and the last two go into 2p orbitals. Since electrons repel each other, the electrons in the 2p orbitals go into separate orbitals before pairing up. Hund’s rule is a statement of this principle, stating that no electrons are paired in a given orbital until all the orbitals of the same sub-level have received at least one electron. If you look at the orbital representation for the electron configuration of oxygen, shown below, you will see that all three 2p orbitals have at least one electron and only one 2p orbital has two electrons.



The image below shows an example of the orbital representation for zinc.



# The Electron Configuration Code

A shorthand method for representing the electron arrangement in the orbital representation is called the electron configuration code. The **electron configuration code** lists the number of the principal energy level followed by the letter of the sub-level type. A superscript is placed on the sub-level letter to indicate the number of electrons in that sub-level. For example, we looked at the orbital representation of carbon in the previous section. The electron configuration code for that configuration is  $1s^22s^22p^2$ . This code shows that the  $1s$  sub-level holds two electrons, the  $2s$  sub-level holds two electrons, and the  $2p$  sub-level holds two electrons. The electron configuration is  $1s^22s^22p^4$  for oxygen and  $1s^22s^22p^63s^23p^64s^23d^{10}$  for zinc. The electron configuration code follows the same filling order as the orbital representation, so the  $4s$  orbital comes before the  $3d$  orbital.

The chart below shows the electron configuration code and the orbital representation for the first seven elements in the periodic table.

Element Symbol	Electron Configuration Code	Orbital Representation
H	$1s^1$	1 $\uparrow$ s
He	$1s^2$	1 $\uparrow\downarrow$ s
Li	$1s^22s^1$	2 $\uparrow$ $\circ$ $\circ$ $\circ$ s                      p 1 $\uparrow\downarrow$ s
Be	$1s^22s^2$	2 $\uparrow\downarrow$ $\circ$ $\circ$ $\circ$ s                      p 1 $\uparrow\downarrow$ s
B	$1s^22s^22p^1$	2 $\uparrow\downarrow$ $\uparrow$ $\circ$ $\circ$ s                      p 1 $\uparrow\downarrow$ s
C	$1s^22s^22p^2$	2 $\uparrow\downarrow$ $\uparrow$ $\uparrow$ $\circ$ s                      p 1 $\uparrow\downarrow$ s
N	$1s^22s^22p^3$	2 $\uparrow\downarrow$ $\uparrow$ $\uparrow$ $\uparrow$ s                      p 1 $\uparrow\downarrow$ s

As the electron configurations get longer and longer, it becomes tedious to write them out. A shortcut has been devised to make this process less tedious. Consider the electron configuration for potassium:  $1s^22s^22p^63s^23p^64s^1$ . The electron configuration for argon is similar, except that it has one less electron than potassium:  $1s^22s^22p^63s^23p^6$ . It is acceptable to use [Ar] to represent the electron configuration for argon and [Ar] $4s^1$  to represent the code for potassium. Using this shortcut, the electron configuration for calcium would be [Ar] $4s^2$ , and the code for scandium would be [Ar] $4s^23d^1$ . Note that generally only noble gases (the last column on the periodic table) have their electron configurations abbreviated in this form. For example, aluminum would rarely (if ever) be represented as [Na] $3p^1$ . Instead, the electron configuration would be written as [Ne] $3s^23p^1$ .



# Quantum Numbers

Recall from the chapter on “The Quantum Mechanical Model of the Atom” that there are four quantum numbers coming from the solutions and extension to Schrödinger’s equation. These numbers have a precise correspondence to the orbital representation discussed in the previous section. The principal quantum number,  $n$ , is the number of the energy level and is sometimes referred to as the electron shell. This number may be any non-zero positive integer (1, 2, 3, . . .). Although there is no theoretical maximum principal quantum number, in practice we will not need a number greater than 7.

The second quantum number is called the **angular momentum quantum number** and is designated by the script letter  $\ell$ . The angular momentum quantum number describes the sub-shell or region in which the electron can be found. This quantum number dictates orbital space and may have any integer value from 0 to  $n - 1$ . Therefore, if  $n = 1$ ,  $\ell$  may only have the value 0. Possible  $\ell$  values for the first four energy levels are shown in **Table 1.1**. The  $\ell$  quantum number corresponds to the sub-energy levels in our orbital representation. The  $s$  sub-level is represented by  $\ell = 0$ , the  $p$  sub-level by  $\ell = 1$ , the  $d$  sub-level by  $\ell = 2$ , and the  $f$  sub-level by  $\ell = 3$ . Theoretically, there are more sub-levels, but these four are enough to describe the ground state of all known elements.

Table 1.1: The First Two Quantum Numbers

$n$ value	Possible $\ell$ values	Sub-energy level
1	0	$0 = s$
2	0, 1	$0 = s, 1 = p$
3	0, 1, 2	$0 = s, 1 = p, 2 = d$
4	0, 1, 2, 3	$0 = s, 1 = p, 2 = d, 3 = f$

Therefore, the first two quantum numbers for an electron tell us what energy level and what sub-level the electron occupies.

The third quantum number is called the **magnetic quantum number** and is designated by the letter  $m$  or  $m_\ell$ . The magnetic quantum number may have integer values from  $\ell$  to  $-\ell$ , as seen in **Table 1.2**. This quantum number describes the orientation of a particular orbital in space.

Table 1.2: The Third Quantum Number

$\ell$ value	Possible $m_\ell$ values
$\ell = 0$	$m_\ell = 0$
$\ell = 1$	$m_\ell = -1, 0, +1$
$\ell = 2$	$m_\ell = -2, -1, 0, +1, +2$
$\ell = 3$	$m_\ell = -3, -2, -1, 0, +1, +2, +3$

The  $m_\ell$  quantum number corresponds to the individual orbitals within a sub-energy level. When  $\ell = 0$ , the sub-energy level is  $s$ . Since sub-level  $s$  has only one orbital,  $m = 0$  refers to that single orbital. When  $\ell = 1$ , the sub-energy level is  $p$ , which has three orbitals. Each of the  $m$  numbers (-1, 0, +1) refer to one of the  $p$  orbitals. Similarly, the five possible  $m_\ell$  values for  $\ell = 2$  correspond to the five possible  $d$  orbitals, and the seven possible values for  $\ell = 3$  correspond to the seven  $f$  orbitals.

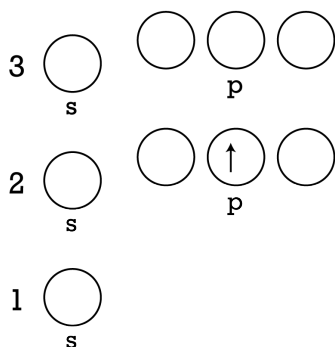
The final quantum number is called the **spin quantum number** and is designated by the letter  $s$  (not to be confused with the  $s$  sub-level, it is sometimes denoted  $m_s$  in other sources). The only possible values for  $s$  are either  $+\frac{1}{2}$  or  $-\frac{1}{2}$ . Electron “spin” is a property that is not encountered in the macroscopic world, but

it has characteristics that are similar to the angular momentum of a charged, spinning sphere. In order for two electrons to occupy the same orbital, they must have different spin values. The first three quantum numbers are specified by the orbital, so adding two electrons with the same spin would violate the Pauli Exclusion Principle.

A complete set of four quantum numbers will identify any individual electron in the atom. For example, suppose we have the quantum numbers 2, 1, 1,  $+\frac{1}{2}$  for an electron. This electron will be in the second energy level, the  $p$  sub-level, the  $+1$  orbital, and it will be spinning in the  $+\frac{1}{2}$  direction. This set of four numbers specifically identifies one particular electron in the electron cloud, so no other electron may have the same combination of numbers.

### Example:

Give four possible quantum numbers for the electron indicated in the orbital representation shown below.



The electron is in the second energy level, hence  $n = 2$ . It is in the  $p$  sub-energy level, so we can let  $\ell = 1$ . It is in the middle orbital, which we can assign as  $m = 0$ . It is pointing up, which we can assign as  $s = +\frac{1}{2}$ .

We should take a moment to recognize that the three orbitals in the  $p$  sub-levels are sometimes designated as the  $p_x$ ,  $p_y$ , and  $p_z$  orbitals. This designation points out that the spatial orientation of the three orbitals corresponds to the three dimensions in a three-dimensional coordinate system. Since an atom can be rotated through any angle, designating a specific orbital to be called  $p_x$ ,  $p_y$ , or  $p_z$  becomes pointless. As long as we recognize there are three mutually perpendicular orbitals, it doesn't make any difference which one is called  $p_x$ . This same reasoning applies to assigning  $+\frac{1}{2}$  and  $-\frac{1}{2}$  to the  $s$  quantum numbers. We must recognize that two electrons may occupy the same orbital and must have a different  $s$  quantum value, but which electron is assigned to be  $+\frac{1}{2}$  and which one to be  $-\frac{1}{2}$  makes no difference.

This video discusses the physical interpretation of quantum numbers (1g): <http://www.youtube.com/watch?v=e9N2h8c6dE4> (6:22).



Figure 1.1: ([Watch Youtube Video](#))

<http://www.ck12.org/flexbook/embed/view/368>

The other video explains the four quantum numbers which give the "address" for an electron in an atom

(1g): <http://www.youtube.com/watch?v=63u7A2NIiyU> (2:54).



Figure 1.2: ([Watch Youtube Video](#))  
<http://www.ck12.org/flexbook/embed/view/369>

## Lesson Summary

- The orbital representation uses circles or lines to represent each atomic orbital.
- In the orbital representation, the distance from the bottom of the chart indicates the energy of each energy level and sub-level.
- The Aufbau principle states that as electrons are added to “build up” the elements, each electron is placed in the lowest energy orbital available.
- Hund’s rule states that no electrons are paired in a given orbital until all the orbitals of the same sub-level have received at least one electron.
- A shorthand method of representing electron arrangement is called the electron configuration code.
- The electron configuration code lists the number of the principal energy level followed by the letter of the sub-level type. A superscript is placed on the sub-level letter to indicate the number of electrons in that sub-level.
- A set of four quantum numbers specifically identifies one particular electron in the electron cloud, and no two electrons may have the same four quantum numbers.
- The principal quantum number is the number of the energy level.
- The second quantum number is called the angular momentum quantum number. This number indicates the subshell in which an electron can be found.
- The third quantum number is called the magnetic quantum number and describes the orientation in space of a particular orbital.
- The final quantum number is called the spin quantum number and indicates the orientation of the angular momentum of an electron in an atom.

## Further Reading / Supplemental Links

These two websites examine the discovery of the spin quantum number.

- <http://www.lorentz.leidenuniv.nl/history/spin/goudsmit.html>
- [http://www.ethbib.ethz.ch/exhibit/pauli/elektronenspin\\_e.html](http://www.ethbib.ethz.ch/exhibit/pauli/elektronenspin_e.html)

This video provides an introduction to the electron configuration of atoms.

- <http://www.youtube.com/watch?v=fv-YeI4hcQ4>

## Review Questions

1. How many energy levels are used for known atoms?
2. Which principal energy level holds a maximum of eight electrons?
3. Which sub-energy level holds a maximum of six electrons?
4. Which sub-energy level holds a maximum of ten electrons?
5. Compare the energy of an electron in the  $3d$  sub-level to that of an electron in the  $4s$  sub-level of the same atom. Which one is greater in energy?
6. When the  $3p$  sub-level of an energy level is filled, where does the next electron go?
7. If all the orbitals in the first two principal energy levels are filled, how many electrons are required?
8. How many electrons are in the electron cloud of a neutral carbon atom?
9. In which principal energy level and sub-level of the carbon atom is the outermost electron located?
10. How many electrons are in the  $2p$  sub-energy level of a neutral nitrogen atom?
11. Which element's neutral atoms will have the electron configuration  $1s^2 2s^2 2p^6 3s^2 3p^1$ ?
12. What energy level and sub-level immediately follow  $5s$  in the filling order?
13. When the  $4f$  and all lower sub-energy levels are filled, what is the total number of electrons in the electron cloud?
14. What is the outermost energy level and sub-level used in the electron configuration of potassium?
15. How many electrons are present in the  $2p$  orbitals of a neutral fluorine atom?
16. Which element will have the electron configuration code  $1s^2 2s^2 2p^6 3s^2 3p^2$ ?
17. Write the electron configuration code for chlorine.
18. In an atom, what is the maximum number of electrons that can have the principal quantum number  $n = 3$ ?
19. In an atom, what is the maximum number of electrons that can have the quantum numbers  $n = 3$ ,  $\ell = 2$ ?
20. In an atom, what is the maximum number of electrons that can have the quantum numbers  $n = 3$ ,  $\ell = 2$ ,  $m_\ell = 0$ ?
21. The following set of quantum numbers is not possible:  $n = 1$ ,  $\ell = 1$ ,  $m_\ell = 0$ . Explain why not.
22. The following set of quantum numbers is not possible:  $n = 2$ ,  $\ell = -1$ ,  $m_\ell = 0$ . Explain why not.

## 1.2 Valence Electrons

### Lesson Objectives

The student will:

- define valence electrons.
- indicate the number of valence electrons for selected atoms.
- draw Lewis dot diagrams for selected atoms.

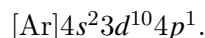
### Vocabulary

- core electron
- Lewis dot diagram
- valence electron
- valence shell

## Introduction

The electrons in the outermost shell are the **valence electrons**. These are the electrons in an atom that can participate in a chemical reaction. Since filled *d* or *f* subshells are seldom disturbed in a chemical reaction, we can define valence electrons as follows: The electrons in an atom that are not present in the previous noble gas (elements in the last column of the periodic table), ignoring filled *d* or *f* subshells.

Gallium has the following electron configuration:



The 4*s* and 4*p* electrons can be lost in a chemical reaction, but the electrons in the filled 3*d* subshell cannot. Gallium therefore has three valence electrons.

## Determining Valence Electrons

Valence electrons are the electrons in the outermost principal quantum level of an atom. Sometimes, the outermost energy level is called the **valence shell**. The electrons that are not in the outermost energy level are known as **core electrons**. The number of valence electrons for an atom can be seen in either the orbital representation or the electron configuration. The electron configuration for magnesium is  $1s^22s^22p^63s^2$ . The outer energy level for this atom is  $n = 3$ , and it has two electrons in this energy level. Therefore, magnesium has two valence electrons. The electron configuration for sulfur is  $1s^22s^22p^63s^23p^4$ . The outer energy level in this atom is  $n = 3$ . It holds six electrons, so sulfur has six valence electrons. The electron configuration for gallium is  $1s^22s^22p^63s^23p^64s^23d^{10}4p^1$ . The outer energy level for this atom is  $n = 4$ , and it contains three electrons. You must recognize that even though the 3*d* sub-level is mixed in among the 4*s* and 4*p* sub-levels, 3*d* is *not* in the outer energy level, so the electrons in the 3*d* sub-level are *not* valence electrons. Gallium has three electrons in the outer energy level and therefore has three valence electrons. The identification of valence electrons is vital because the chemical behavior of an element is determined primarily by the arrangement of the electrons in the valence shell.

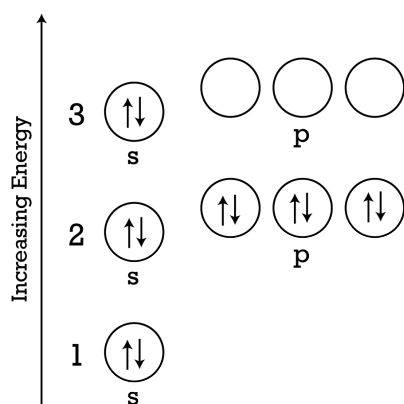
## Lewis Dot Diagrams

We can represent the valence electrons of an atom using a **Lewis dot diagram**. A Lewis dot diagram indicates valence electrons as dots around the symbol of the element. A Lewis dot diagram, also known as an electron dot diagram, consists of two parts: the elemental symbol, which represents the core (inner electrons and nucleus), and a group of dots. There can be between 1 through 8 dots, with each dot representing a valence electron.

The electron configuration for sodium (Na) is:  $1s^22s^22p^63s^1$ . The core electrons are  $1s^22s^22p^6$ . The valence electron is  $3s^1$ . To draw a Lewis electron dot diagram for sodium, you can picture the symbol for sodium in a box with four sides. Each side of the box represents either the *s* or one of the three *p* orbitals in the outermost energy level of the atom. The first and second valence electrons are placed on the side representing the *s* orbital, and the next electrons are placed in the *p* orbitals. The electron dot diagram also follows Hund's rule, so all the *p* orbitals in the diagram must contain at least one electron before any of the *p* orbitals can have a second electron. Since each side can only have two electrons, there can be a maximum of  $2 \times 4 = 8$  electrons on any Lewis electron dot diagram. For sodium, a single dot will represent its one valence electron. Thus, the Lewis dot formula for sodium is:



Look at the electron configuration for magnesium shown below. Magnesium is the twelfth element on the periodic table.

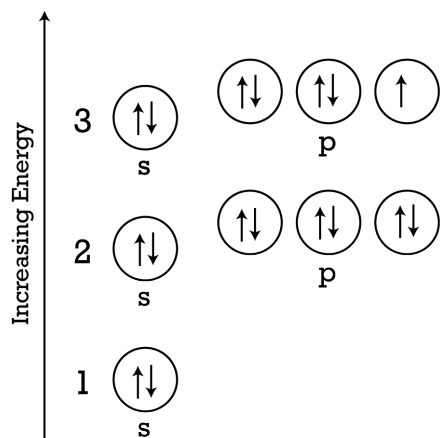


Electron Configuration for Magnesium (Mg)

To draw the Lewis electron dot diagram, picture the symbol for magnesium in a box, which represents all of the core electrons ( $1s^2 2s^2 2p^6$ ). Then, place the valence electrons around the sides of the box with each side representing an orbital in the outermost energy level. How many valence electrons does magnesium have? There are 2 valence electrons, as indicated by  $3s^2$ . Therefore, the Lewis electron dot formula for magnesium is:

**Mg:**

Look at the electron configuration for chlorine.



Electron Configuration for Chlorine (Cl)

The electron configuration for chlorine (Cl) could be written as:  $1s^2 2s^2 2p^6 3s^2 3p^5$ . The core electrons would be  $1s^2 2s^2 2p^6$ , while the valence electrons would be in the third shell (or where  $n = 3$ ). Therefore, chlorine has 7 valence electrons. The Lewis electron dot diagram would look like the following:



**Example:**

Write the Lewis electron dot formula for:

1. Oxygen
2. Sulfur
3. Potassium
4. Carbon

### Solution:

1. Oxygen has the electron configuration:  $1s^2 2s^2 2p^4$

The Lewis electron dot formula is:



2. Sulfur has the electron configuration:  $1s^2 2s^2 2p^6 3s^2 3p^4$

The Lewis electron dot formula is:



3. Potassium has the electron configuration:  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

The Lewis electron dot formula is:



4. Carbon has the electron configuration:  $1s^2 2s^2 2p^2$

The Lewis electron dot formula is:

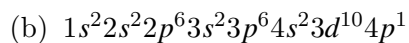


## Lesson Summary

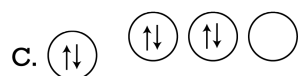
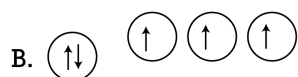
- Valence electrons are the electrons in the outermost principal quantum level of an atom.
- The identification of valence electrons is vital because the chemical behavior of an element is determined primarily by the arrangement of the electrons in the valence shell.
- In a Lewis dot diagram, the chemical symbol is surrounded by a number of dots. Each dot represents a valence electron.

## Review Questions

1. How many valence electrons are present in the following electron configuration code?  
(a)  $1s^2 2s^2 2p^6 3s^2 3p^3$



- How many valence electrons does magnesium have?
- How many valence electrons does oxygen have?
- Which of the following is the correct orbital representation for the valence shell of phosphorus?



D. None of these.

- What is the maximum number of dots in a Lewis dot diagram?
- Draw the Lewis dot diagram for lithium.
- Draw the Lewis dot diagram for calcium.
- Draw the Lewis dot diagram for bromine.
- Draw the Lewis dot diagram for selenium.

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