Electron Configurations, Orbital Notations and Quantum Numbers
Understanding Electron Arrangement and Oxidation States

OBJECTIVE
Students will learn to write correct electron configurations, orbital notations and quantum numbers for the valence electron in certain elements. Students will learn to justify oxidation states based on electron configurations which will enhance their understanding of chemical formula writing.

LEVEL
Chemistry

NATIONAL STANDARDS
UCP.1, UCP.2, B.1, B.2

CONNECTIONS TO AP
AP Chemistry:
I. Structure of Matter  A. Atomic theory and atomic structure  4. Electron energy levels: atomic spectra, quantum numbers, atomic orbitals

TIME FRAME
90 minutes

MATERIALS FOR THE TEACHER
Transparency of Periodic Table [unless one hangs on your classroom wall]
Transparency of the Diagonal Rule

TEACHER NOTES
This lesson should be taught after students have mastered basic atomic structure and layout of the periodic table. This lesson will introduce skills necessary for formula writing and nomenclature units. Writing electron configurations, orbital notations and quantum numbers are fundamental to a first-year foundational course. All serve to help students understand the logic of writing chemical formulas and predicting oxidation states.
Suggested Teaching Procedure:
1. Before presenting to students, complete the student activity yourself. You will need a copy of the periodic table included with this activity.

2. Once you have mastered the skill, introduce the lesson by demonstrating your predicting prowess to engage the students’ curiosity.

3. Keep the following points in mind during your introduction:
   - The periodic table is arranged according to electron configurations. With practice, you’ll be able to give the ending term of the electron configuration within a few seconds.
   - Study the periodic table below that shows the patterns of the electron configurations.
   - Do not share this periodic table with the students, keep a copy handy as you impress them with your knowledge of electron configurations and engage their curiosity. Hopefully, by the end of this activity, they will deduce these patterns for themselves.

   Period numbers represent energy levels.

   Notice that there are only two elements in the first period, which represents the first energy level. The second period is separated so that two elements are together on the left and six are together on the right. The second energy level has two sublevels s and p; s-sublevels contain two electrons when full while p-sublevels contain six electrons when full. The s-block of the table contains two columns of elements grouped together on the far left. These elements are the alkali metals and the alkaline earth metals, groups IA & IIA. The p-block consists of the six columns of elements on the far right of the table, groups IIIA–VIIIA.

   - Notice the third period looks like the second.
   - The third energy level is the first energy level with the capacity to contain a d-sublevel. You might expect to see the insertion of the d-sublevel here, but it is delayed until the 4th period.
• The delay is because of the interference between electrons. The fact is that it takes less energy for an electron to be placed in the 4s sublevel than in the 3d; as a result, the 4s fills before the 3d. Therefore, on the table the 10 electrons representing 3d are placed immediately after the 4s and begin the section called the transition metals. Students often mis-number the d-block elements. You must remind them that the first time you get a d-sublevel at all, it is in the third energy level so that top row of d-elements must be 3d while the one below must be 4d, etc.

• Finally, look at the rare earth elements at the bottom of the table. There are 14 of them which correspond to the filling of the f sublevels. There are only 2 rows in the f-block; the top row must be 4f since that’s the first energy level with an f-sublevel and the bottom row must be 5f.

_A practice exercise:_
Using the periodic table, determine the electron configuration for the valence electron in sulfur.

• First locate sulfur on the periodic table; it is in the 3rd period in the p-block of elements.

• Count from left to right within the p-block and you determine the valence electron has an ending configuration of 3p^4_. The Aufbau principle states that all lower energy sublevels must be full, so its entire electron configuration is 1s^22s^22p^63s^23p^4.

• The noble gas short-hand allows us to substitute the noble gas symbol for the core electrons. In the case of sulfur it would be [Ne] 3s^23p^4. It is up to you whether or not you want students to short-hand the electron configurations on their answer page.

4. Start class by introducing the student handout. An example for determining the electron configuration for sulfur is given in the student directions. The electron configuration and orbital notation for oxygen are given as an example in the “time to get the lingo straight” section.

5. This is where you show off your skills—remember to keep your periodic table with the sublevels labeled handy for moments of panic. Ask students to practice by writing the electron configurations for the following elements in class before they begin their answer sheet exercise. As you announce each element, also announce your prediction for the ending electron configuration—do this by using the element’s position on the periodic table. As you gain confidence, let the students pick their favorite elements and try to stump you. Do not divulge your secrets. Perhaps offer extra credit to the first student that can explain the source of your wisdom to the class. If challenged, students will deduce this pattern on their own.

<table>
<thead>
<tr>
<th>Element</th>
<th>ENDING Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
<td>3s^1</td>
</tr>
<tr>
<td>Ni</td>
<td>3d^8</td>
</tr>
<tr>
<td>C</td>
<td>2p^2</td>
</tr>
<tr>
<td>I</td>
<td>5p^3</td>
</tr>
<tr>
<td>Y</td>
<td>4d^1</td>
</tr>
<tr>
<td>Es*</td>
<td>5f^1</td>
</tr>
</tbody>
</table>

* Counting can be tricky with the Rare Earth elements. The periodic table presented to the students places La and Ac as f^1, so Es would be f^{11}.
ANSWERS TO THE CONCLUSION QUESTIONS

**ANALYSIS**

<table>
<thead>
<tr>
<th>Question Number</th>
<th>Element</th>
<th>Electron Configuration</th>
<th>Non Core Orbital Notation</th>
<th>Set of Quantum Numbers for the LAST Electron to Fill</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td>K</td>
<td>[Ar] 4s(^1)</td>
<td>↑ 4s</td>
<td>4, 0, 0, +½</td>
</tr>
<tr>
<td>2.</td>
<td>Fe</td>
<td>[Ar] 4s(^2)3d(^6)</td>
<td>↑↑ 4s 3d</td>
<td>3, 2, −2, −½</td>
</tr>
<tr>
<td>3.</td>
<td>N</td>
<td>1s(^2)2s(^2)2p(^3)</td>
<td>↑↑ ↑↑ 2p</td>
<td>2, 1, 1, +½</td>
</tr>
<tr>
<td>4.</td>
<td>Sn</td>
<td>[Kr] 5s(^2) 4d(^{10})5p(^2)</td>
<td>↑↑ 5s 4d</td>
<td>5, 1, 0, +½</td>
</tr>
<tr>
<td>5.</td>
<td>Br</td>
<td>[Ar] 4s(^2) 3d(^{10})4p(^5)</td>
<td>↑↑ 4s 3d</td>
<td>4, 1, 0, −½</td>
</tr>
<tr>
<td>6.</td>
<td>Ba</td>
<td>[Xe] 6s(^2)</td>
<td>↑ 6s</td>
<td>6, 0, 0, −½</td>
</tr>
<tr>
<td>7.</td>
<td>Ni</td>
<td>[Ar] 4s(^2)3d(^8)</td>
<td>↑↑ 4s 3d</td>
<td>3, 2, 0, −½</td>
</tr>
<tr>
<td>8.</td>
<td>P</td>
<td>[Ne] 3s(^2) 3p(^3)</td>
<td>↑↑ 3p</td>
<td>3, 1, 1, +½</td>
</tr>
<tr>
<td>9.</td>
<td>Zr</td>
<td>[Kr] 5s(^2) 4d(^2)</td>
<td>↑↑ 5s 4d</td>
<td>4, 2, −1, +½</td>
</tr>
<tr>
<td>10.</td>
<td>U*</td>
<td>[Rn] 7s(^2) 5f(^4)</td>
<td>↑↑ 7s 5f</td>
<td>5, 3, 0, +½</td>
</tr>
<tr>
<td>11.</td>
<td>Ag**</td>
<td>[Kr] 5s(^2) 4d(^9)</td>
<td>↑↑ 5s 4d</td>
<td>4, 2, 1, −½</td>
</tr>
<tr>
<td>12.</td>
<td>Mg</td>
<td>1s(^2) 2s(^2)2p(^6)3s(^2)</td>
<td>↑↑ 3s</td>
<td>3, 0, 0, −½</td>
</tr>
<tr>
<td>13.</td>
<td>Kr</td>
<td>[Ar] 4s(^2) 3d(^{10})4p(^6)</td>
<td>↑↑ 4s 3d</td>
<td>4, 1, 1, −½</td>
</tr>
<tr>
<td>14.</td>
<td>As</td>
<td>[Ar] 4s(^2) 3d(^{10})4p(^3)</td>
<td>↑↑ 4s 3d</td>
<td>4, 1, 1, +½</td>
</tr>
<tr>
<td>15.</td>
<td>W</td>
<td>[Xe] 6s(^2)4f(^{14})5d(^4)</td>
<td>↑↑ 6s 5d</td>
<td>5, 2, 1, +½</td>
</tr>
<tr>
<td>16.</td>
<td>Fr</td>
<td>[Rn] 7s(^1)</td>
<td>↑ 7s</td>
<td>7, 0, 0, +½</td>
</tr>
<tr>
<td>17.</td>
<td>Pu*</td>
<td>[Rn] 7s(^2) 5f(^6)</td>
<td>↑↑ 7s 5f</td>
<td>5, 3, 2, +½</td>
</tr>
<tr>
<td>18.</td>
<td>B</td>
<td>1s(^2)2s(^2)2p(^1)</td>
<td>↑↑ 2p</td>
<td>2, 1, −1, +½</td>
</tr>
<tr>
<td>19.</td>
<td>Mn</td>
<td>[Ar] 4s(^2) 3d(^5)</td>
<td>↑↑ 4s 3d</td>
<td>3, 2, 2, +½</td>
</tr>
<tr>
<td>20.</td>
<td>I</td>
<td>[Kr] 5s(^2) 4d(^{10})5p(^5)</td>
<td>↑↑ 5s 3d</td>
<td>5, 1, 0, −½</td>
</tr>
</tbody>
</table>

**Students may also write [Kr] 5s\(^1\) 4d\(^{10}\) which is how it really exists in which case its quantum number set would be 4, 2, 2, −½.**
Some books will list the rare earth metals as having an electron in the d-sublevel before beginning to fill the f-sublevel. This will reduce the number of electrons shown in the table above by one, but the one electron placed in the previous d-sublevel should be shown in order for credit to be given. You will find that different periodic tables break the f-block differently. Some place the break before La, others after La. Some have 14 boxes for the f-block at the bottom of the table, some have 15.

CONCLUSION QUESTIONS

1. Iron has two common oxidation states, +2 and +3. Justify each of these oxidation states. Draw the orbital notation of the neutral atom and each oxidation state as part of your justification.
   - Neutral atom: $\uparrow\downarrow$ 4s  \[\uparrow\uparrow \uparrow \uparrow \uparrow \uparrow\] 3d
   - +2 oxidation state — 2 electrons are lost and they come from the outermost 4s leaving: $\uparrow\downarrow$ 4s  \[\uparrow\uparrow \uparrow \uparrow \uparrow \uparrow\] 3d
   - +3 oxidation state — 3 electrons are lost. Losing one of the d-electrons minimized electron-electron repulsions: $\uparrow$ 4s  \[\uparrow\uparrow \uparrow \uparrow \uparrow \uparrow\] 3d

2. Nitrogen has a common oxidation state of –3. Justify this oxidation state. Draw the orbital notation for the neutral atom and oxidation state as part of your justification.
   - Neutral atom: $\uparrow\downarrow$ 1s  \[\uparrow\downarrow \uparrow \uparrow \uparrow \uparrow \uparrow\] 2s  \[\uparrow\downarrow \uparrow \uparrow \uparrow \uparrow \uparrow\] 2p
   - −3 oxidation state, three electrons are gained forming a stable octet: $\uparrow\downarrow$ 1s  \[\uparrow\downarrow \uparrow \uparrow \uparrow \uparrow \uparrow\] 2s  \[\uparrow\downarrow \uparrow \uparrow \uparrow \uparrow \uparrow\] 2p

3. Silver has only one oxidation state, +1. Justify this oxidation state even though most transition metals have an oxidation state of +2 among others. Draw the orbital notation for the neutral atom and oxidation state as part of your justification.
   - Neutral atom: $\uparrow\downarrow$ 5s  \[\uparrow\downarrow \uparrow \uparrow \uparrow \uparrow \uparrow\] 4d
   - +1 oxidation state — one electron is lost. This electron will come from the outermost sublevel which is 5s. The movement of the remaining lone electron further stabilizes the ion: $\uparrow\downarrow$ 5s  \[\uparrow\downarrow \uparrow \uparrow \uparrow \uparrow \uparrow\] 4d

There is only a slight energy difference between s and d electrons. The further from the nucleus, the less this difference becomes. It is not uncommon to see electrons shift to fill the d-orbitals.
4. Manganese has a common oxidation state of +7. Justify this oxidation state. Draw the orbital notation for the neutral atom and oxidation state as part of your justification.

- Neutral atom: \[
\begin{array}{c}
\uparrow \\
4s \\
\uparrow \uparrow \uparrow \uparrow \uparrow \\
3d \\
\end{array}
\]

- +7 oxidation state — 7 electrons are lost, which are the entire contents of the 4s and 3d orbitals to create the same electron configuration as argon:
\[
\begin{array}{c}
\quad \\
4s \\
\quad \\
\quad \\
3d \\
\quad \\
\end{array}
\]
The Diagonal Rule or Aufbau Series

1s
2s
2p
3s
3p
3d
4s
4p
4d
5s
5p
5d
6s
6p
6d
7s
7p
Electron Configurations, Orbital Notations and Quantum Numbers

Understanding Electron Arrangement and Oxidation States

Chemical properties depend upon the number and arrangement of electrons in an atom. Usually, only the valence or outermost electrons are involved in chemical reactions. The electron cloud is compartmentalized. We model this compartmentalization through the use of electron configurations and orbital notations. The compartmentalization is as follows: energy levels have sublevels which have orbitals within them. We can use an apartment building as an analogy. The atom is the building, the floors of the apartment building are the energy levels, the apartments on a given floor are the orbitals and electrons reside inside the orbitals. There are two governing rules to consider when assigning electron configurations and orbital notations. Along with these rules, you must remember electrons are lazy and they “hate” each other since they find each other quite repulsive, they will fill the lowest energy states first AND electrons repel each other since like charges repel.

Rule 1: The Pauli Exclusion Principle
In 1925, Wolfgang Pauli stated: No two electrons in an atom can have the same set of four quantum numbers. This means no atomic orbital can contain more than TWO electrons and the electrons must be of opposite spin if they are to form a pair within an orbital.

Rule 2: Hund’s Rule
The most stable arrangement of electrons is one with the maximum number of unpaired electrons. It minimizes electron-electron repulsions and thus stabilizes the atom. Here is an analogy. In large families with several children, it is a luxury for each child to have their own room. There is far less fussing and fighting if siblings are not forced to share living quarters. The entire household experiences a lower, less frazzled energy state. Electrons find each other very repulsive, so they too, are in a lower energy state if each “gets their own room” or, in this case, orbital. Electrons will fill an orbital singly, before pairing up in order to minimize electron-electron repulsions. All of the electrons that are single occupants of orbitals have parallel (same direction) spins and are assigned an up arrow. The second electron to enter the orbital, thus forming an electron pair, is assigned a down arrow to represent opposite spin.

PURPOSE
In this activity you will learn to write electron configurations, orbital notations and a set of quantum numbers for the valence electron of selected elements on the periodic table. You will also practice justifying oxidation or valence states using electron configurations and orbital notations.
MATERIALS
Periodic Table found at the end of the student pages

PART I: ELECTRON CONFIGURATIONS
1. Each main energy level has \( n \) sublevels, where \( n \) equals the number of the energy level. That means the first energy level has one sublevel, the second has two, the third has three and so on.

2. The sublevels are named s, p, d, f, g . . . and continue alphabetically. The modern periodic table does not have enough elements to necessitate the use of sublevels beyond f. Why s, p, d, f? Early on in the development of this model, the names of the sublevels came from sharp, principle, diffuse and fundamental, words used in describing the spectral lines of hydrogen.

3. It may be easier for you to understand this by studying the table presented below:

<table>
<thead>
<tr>
<th>Energy Level</th>
<th>Number of Sublevels</th>
<th>Names of Sublevels</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1</td>
<td>s</td>
</tr>
<tr>
<td>2</td>
<td>2</td>
<td>s, p</td>
</tr>
<tr>
<td>3</td>
<td>3</td>
<td>s, p, d</td>
</tr>
<tr>
<td>4</td>
<td>4</td>
<td>s, p, d, f</td>
</tr>
<tr>
<td>5</td>
<td>5</td>
<td>s, p, d, f, g</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Sublevel Name</th>
<th>s</th>
<th>p</th>
<th>d</th>
<th>f</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number of orbitals</td>
<td>1</td>
<td>3</td>
<td>5</td>
<td>7</td>
</tr>
<tr>
<td>Maximum number of electrons</td>
<td>2</td>
<td>6</td>
<td>10</td>
<td>14</td>
</tr>
</tbody>
</table>

4. Each sublevel has increasing odd numbers of orbitals available. \( s = 1, \ p = 3, \ d = 5, \ f = 7. \)

Each orbital can hold only two electrons and they must be of opposite spin. An s-sublevel holds 2 electrons, a p-sublevel holds 6 electrons, a d-sublevel holds 10 electrons, and an f-sublevel holds 14 electrons.
5. The filling of the orbitals is related to energy. Remember, electrons are lazy, much like us! Just as you would place objects on a bottom shelf in an empty store room rather than climb a ladder to place them on a top shelf, expending more energy—electrons fill the lowest sublevel available to them. Use the diagonal rule as your map when determining electron configurations for any of the elements.

Using the diagonal rule you can quickly determine the electron configuration for the outermost valence electron in sulfur. First locate sulfur on the periodic table and notice that the atomic number of sulfur is 16. That means it has 16 protons and 16 electrons in a neutral atom. The first two electrons go into the 1s sublevel and fill it, the next two go into the 2s sublevel and fill it. That leaves 12 more electrons to place. The next six go into the 2p sublevel, filling it and leaving six more. Two of them go into the 3s sublevel, filling it and the remaining four go into the 3p sublevel. The completed electron configuration looks like this: 1s²2s²2p⁶3s²3p⁴.

6. Complete the electron configuration column for the rows in which element symbols are given on your student answer page.

PART II: ORBITAL NOTATIONS
Orbital notation is a drawing of the electron configuration. It is very useful in determining electron pairing and thus predicting oxidation numbers. The orbital notation for sulfur would be represented as follows:

```
1 2 3 4 5 8 6 9 7 10 11 12 13 16 14 15
↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓
1s 2s 2p 3s 3p
```

The electrons are numbered as to the filling order. Notice electrons 5, 6, and 7 went into their own orbitals before electrons 8, 9, and 10 forced a pairing to fill the 2p sublevel. This is an application of Hund’s rule which minimizes electron-electron repulsions. The same filling order is repeated in the 3p sublevel.

It’s time to get the lingo straight!
**Electron Configurations**

Group the 1’s, 2’s, etc. TOGETHER and it looks like this:

\[ 1s^2 2s^2 2p^4 \]

Which element has this electron configuration?

**Orbital notations**

Use blanks to represent orbitals and arrows to represent electrons and it looks like this:

\[
\begin{array}{ccccccc}
1 & 2 & 3 & 4 & 5 & 8 & 6 & 7 \\
\uparrow \uparrow & \uparrow \uparrow & \uparrow \uparrow & \uparrow \uparrow & \uparrow \uparrow & \uparrow \uparrow & \uparrow \uparrow \\
1s & 2s & 2p & & & & &
\end{array}
\]

The electrons are numbered as to the filling order. Notice electrons 5, 6, 7 went into their own orbitals before electron 8 forced a pairing. This minimizes repulsion.

Which element has this orbital notation?

1. Complete the orbital notation column for the elements listed on your student answer page. Also use the orbital notations given to complete both the Element column and the Electron configuration column on your student answer page.

**PART III: JUSTIFYING OXIDATION STATES**

Elements in compounds have oxidation states. These oxidation states determine an element’s behavior in the company of other elements. Your understanding of oxidation states will become very important as you learn to write correct chemical formulas for compounds. Some elements have only one oxidation state, while others have several. In general, the representative elements, those groups or families numbered as IA – VIIIA, have the oxidation states listed on the periodic table below.

The transition metals generally have several oxidation states possible.

Learn the following; it will help you make your predictions:
- Metals (found to the left of the stair-step line) lose electrons to either minimize electron-electron repulsions or eliminate their valence electrons entirely.
- Nonmetals tend to gain electrons to acquire an octet of electrons. An octet means the atom has eight valence electrons arranged as \( ns^2 np^6 \) where \( n \) corresponds to the main or principal energy level.
• Transition metals generally have an oxidation state of +2 since they lose the s² that was filled just before the d-sublevel began filling.
• Electrons in the d-sublevels are very similar in energy to those in the s-sublevel preceding them. This means that 3d electrons are similar in energy to 4s electrons and 4d are similar to 5s, etc.
• Noble gases have an octet naturally, so they generally do not react.

Let’s practice.
• Sulfur has many oxidation states. Use an orbital notation to justify its most common −2 oxidation state:

Sulfur has a valence electron configuration of 3s²3p⁴. Start by drawing its orbital notation for the outermost, valence electrons.

\[
\begin{array}{c}
\text{[Ne]} \\
\uparrow \downarrow \\
\uparrow \downarrow \\
\uparrow \\
\end{array} \\
3s \\
3p
\]

Sulfur is a nonmetal and tends to gain electrons, creating the −2 charge. Gaining two electrons gives it an octet of 3s²3p⁶.

• Copper has two common oxidation states, +1 and +2. Justify both oxidation states:
Copper has an ending electron configuration of 4s²3d⁹. Start by drawing its orbital notation for the outermost, valence electrons.

\[
\begin{array}{c}
\text{[Ar]} \\
\uparrow \downarrow \\
\uparrow \downarrow \\
\uparrow \downarrow \\
\uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \\
4s \\
3d
\]

Since copper is a transition metal, the +2 oxidation state comes from losing the 4s electrons leaving 4s⁰3d⁹. Almost all of the transition metals lose the s sublevel and have an oxidation state of +2. [Silver is an exception and only makes an oxidation state of +1.] The +1 oxidation state for copper comes from transferring one of the s electrons to the d orbitals to fill that sublevel and then losing the remaining s electron to form 4s⁰3d¹⁰.

1. Complete the conclusion questions that justify oxidation states on your student answer page.
### PART IV: QUANTUM NUMBERS AND ATOMIC ORBITALS

<table>
<thead>
<tr>
<th>Principal quantum number ((n))</th>
<th>Determines the total energy of the electron. Describes the energy level of the electron and refers to the average distance of the electron from the nucleus. (2n^2) electrons may be assigned to an energy level. For (n = 1), 2 electrons. For (n = 2), 8 electrons, etc.</th>
</tr>
</thead>
<tbody>
<tr>
<td>1, 2, 3, 4, 5, etc.</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Angular momentum or azimuthal quantum number ((\ell))</th>
<th>Refers to the sublevels that occur within each principal level and determines the shape of the orbital. Corresponds to the s, p, d, f [in order of increasing energy]. Each (\ell) is a different orbital shape or orbital type. This quantum number has integral values from 0 up to (n-1).</th>
</tr>
</thead>
<tbody>
<tr>
<td>0, 1, 2, 3…</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Magnetic quantum number ((m_\ell))</th>
<th>Specifies which orbital within a sublevel you are likely to find the electron. It determines the orientation of the orbital in space relative to the other orbitals in the atom. This quantum number has values from (-\ell) through zero to (+\ell).</th>
</tr>
</thead>
<tbody>
<tr>
<td>…−2, −1, 0, 1, 2, …</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Spin quantum number ((m_s))</th>
<th>Specifies the value for the spin. Only two possibilities exist: (+\frac{1}{2}) and (−\frac{1}{2}). No more than two electrons can occupy an orbital. In order for two electrons to occupy an orbital, they must have opposite spins.</th>
</tr>
</thead>
<tbody>
<tr>
<td>+ (\frac{1}{2}) or (−\frac{1}{2})</td>
<td></td>
</tr>
</tbody>
</table>

### Determining Quantum Numbers

Now that we know the electron configuration of the valence electron in sulfur is \(3p^4\) based on its position in the periodic table, and we have a picture of how those p-electrons are filling the p-sublevel, the set of quantum numbers for this valence electron are extremely easy to obtain. First, \(n = 3\) since it is a \(3p\) electron. Next it is a \(p\)-electron and p-sublevels have an \(\ell\) value of 1. So far we know 3,1. To get the \(m_\ell\) quantum number we go back to the orbital notation for the valence electron and focus on the \(3p\) sublevel alone. It looks like this:

\[
\begin{array}{cc}
\uparrow & \downarrow \\
-1 & 0 \\
\uparrow & \uparrow
\end{array}
\]

Simply number the blanks with a zero assigned to the center blank and increasing negative numbers to the left and increasing positive to the right of the zero. The last electron was number 16 and “landed” in the first blank as a down arrow. This picture gives us the last two quantum numbers of \(m_\ell = -1\) and \(m_s = -\frac{1}{2}\) since it is the second electron to be placed in the orbital.
In summary:

<table>
<thead>
<tr>
<th>Energy Level</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7…</th>
</tr>
</thead>
<tbody>
<tr>
<td># of sublevels</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>5</td>
<td>6</td>
<td>7…</td>
</tr>
<tr>
<td>Names of sublevels</td>
<td>s</td>
<td>s,p</td>
<td>s,p,d</td>
<td>s,p,d,f</td>
<td>s,p,d,f,g</td>
<td>s,p,d,f,g,h</td>
<td>s,p,d,f,g,h,i</td>
</tr>
<tr>
<td>n, principal quantum number</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>5</td>
<td>6</td>
<td>7…</td>
</tr>
<tr>
<td>Name of sublevel</td>
<td>s</td>
<td>p</td>
<td>d</td>
<td>f</td>
<td>g</td>
<td>h</td>
<td>i…</td>
</tr>
<tr>
<td>ℓ, angular momentum quantum number (=[n-1])</td>
<td>0</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>5</td>
<td>6…</td>
</tr>
<tr>
<td># of orbitals (=[-ℓ to +ℓ])</td>
<td>1</td>
<td>3</td>
<td>5</td>
<td>7</td>
<td>9</td>
<td>11</td>
<td>13…</td>
</tr>
<tr>
<td>(m_ℓ) for each orbital within a sublevel</td>
<td>0</td>
<td>−1</td>
<td>0</td>
<td>+1</td>
<td>−2</td>
<td>−1</td>
<td>0</td>
</tr>
</tbody>
</table>

And so on, just pretend you’re in elementary school and make a number line with ZERO in the middle with negative numbers to the left and positive to the right. Make as many blanks as there are orbitals for a given sublevel.

For assigning \(m_σ\), the first electron placed in an orbital [the up arrow] gets the \(+\frac{1}{2}\) and the second one [the down arrow] gets the \(-\frac{1}{2}\).

Try working backwards. Which element has this set of quantum numbers \(5, 1, -1, -\frac{1}{2}\)? First think about the electron configuration. \(n = 5\) and \(ℓ = 1\), so it must be a 5p electron. The \(m_σ\) quantum number corresponds to this orbital notation picture ↓. Be sure and number the blanks and realize that the \(-\frac{1}{2}\) means it is a pairing electron, so the orbital had to be half-filled before pairing could occur, thus for the electron to occupy the \(-1\) position, it must be a \(p^4\) electron. The element has a configuration of \(5p^4\) so it must be Tellurium.

Quantum numbers are a set of 4 numbers that describe an electron’s position within an atom. They are quite easy to determine if you start with the electron configuration. The set of quantum numbers for the \(2p^4\) electron would be \(2, 1, -1, +\frac{1}{2}\). Each electron in the atom has a set of quantum numbers, but you will most often be asked for the set describing the last electron.

1. Write a complete set of four quantum numbers for the LAST electron to fill the electron cloud for the selected elements on your student answer page. Write your answer in the column provided.
# Electron Configurations, Orbital Notations and Quantum Numbers

**Understanding Electron Arrangement and Oxidation States**

## ANALYSIS

Complete this table:

<table>
<thead>
<tr>
<th>Question Number</th>
<th>Element</th>
<th>Electron Configuration</th>
<th>Non Core Orbital Notation [only the outermost orbitals are drawn]</th>
<th>Set of Quantum Numbers for the LAST Non Core Electron to Fill</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td></td>
<td></td>
<td>$\uparrow$ 4s</td>
<td></td>
</tr>
<tr>
<td>2.</td>
<td>Fe</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3.</td>
<td></td>
<td>$1s^22s^22p^3$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4.</td>
<td></td>
<td>$\uparrow\uparrow$ 5s</td>
<td>$\uparrow\uparrow$ 4d $\uparrow\uparrow$ 5p _ _ _ _</td>
<td></td>
</tr>
<tr>
<td>5.</td>
<td>Br</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>6.</td>
<td></td>
<td></td>
<td>$\uparrow\uparrow$ 6s</td>
<td></td>
</tr>
<tr>
<td>7.</td>
<td></td>
<td>[Ar] 4s^23d^4</td>
<td></td>
<td></td>
</tr>
<tr>
<td>8.</td>
<td>P</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>9.</td>
<td></td>
<td></td>
<td>$\uparrow\downarrow$ 5s $\uparrow\uparrow$ _ _ _ _ _ _</td>
<td></td>
</tr>
<tr>
<td>10.</td>
<td>U</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>11.</td>
<td></td>
<td>[Kr] 5s^24d^9</td>
<td></td>
<td></td>
</tr>
<tr>
<td>12.</td>
<td></td>
<td></td>
<td>$\uparrow\downarrow$ 3s</td>
<td></td>
</tr>
<tr>
<td>13.</td>
<td></td>
<td>[Ar] 4s^23d^104p^6</td>
<td></td>
<td></td>
</tr>
<tr>
<td>14.</td>
<td></td>
<td></td>
<td>$\uparrow\downarrow$ 4s $\uparrow\downarrow$ 3d $\uparrow\downarrow$ 4p</td>
<td></td>
</tr>
<tr>
<td>15.</td>
<td>W</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>16.</td>
<td></td>
<td>[Rn] 7s^1</td>
<td></td>
<td></td>
</tr>
<tr>
<td>17.</td>
<td>Pu</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>18.</td>
<td></td>
<td>$1s^22s^22p^1$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>19.</td>
<td></td>
<td></td>
<td>$\uparrow\downarrow$ 4s $\uparrow\downarrow$ _ _ _ _ _ _</td>
<td></td>
</tr>
<tr>
<td>20.</td>
<td>I</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
CONCLUSION QUESTIONS

1. Iron has two common oxidation states, +2 and +3. Justify each of these oxidation states. Draw the orbital notation of the neutral atom and each oxidation state as part of your justification.

2. Nitrogen has a common oxidation state of $-3$. Justify this oxidation state. Draw the orbital notation for the neutral atom and oxidation state as part of your justification.

3. Silver has only one oxidation state, +1. Justify this oxidation state even though most transition metals have an oxidation state of +2 among others. Draw the orbital notation for the neutral atom and oxidation state as part of your justification.

4. Manganese has a common oxidation state of +7. Justify this oxidation state. Draw the orbital notation for the neutral atom and oxidation state as part of your justification.