**ENERGY DIAGRAM OF ORBITALS**

Each line stands for an orbital, which can accommodate a maximum of 2 electrons, of different spins (↑ or ↓)

If there is more than one orbital of the same energy, electrons will spread out into different orbitals before pairing up.

**RULES FOR RENTING APARTMENTS**

1. Rent rooms with lowest rent first.
2. Each room can take up to 2 occupants.
3. The 2 occupants within a room must have different ___.
4. If there are more than 1 room of the same rent, occupants will spread out first before pairing up with roommates.

**RULES FOR FILLING ELECTRONS INTO THE ENERGY DIAGRAM**

1. Electrons fill up levels with the lowest energy first. (**Aufbau Principle**)
2. Each orbital can take up to 2 electrons.
3. The 2 electrons within an orbital must have different spins. (**Pauli Exclusion Principle**)
4. If there are more than 1 orbital of the same energy, electrons will spread out first before pairing up. (**Hund’s Rule**)

The "outershell" is the E level with the highest n that contains electrons.

**Outershell electrons = valence electrons** = e⁻ residing in the outershell. These are the ones that participate in chem. reactions.

The **ground state** of an atom is the state where all the e⁻ are at the lowest E levels possible. There can only be one ground state.

An **excited state** of an atom is a state where an e⁻ is boosted up to a higher E level. There can be more than one excited state. When an atom is in an excited state, it is unstable and will try to return to its ground state.

You should be able to …

1. Fill electrons into E diagrams for each element.
2. Write the electron configuration for each element.
3. Circle the “outershell” or “valence” electrons for each element.
Fill in the electrons for each of the elements, circle the valence electrons and then write the electron configuration.

\[ ^{23}_{\text{V}} \] \[ ^{33}_{\text{As}} \] \[ ^{24}_{\text{Cr}} \) (an exception you must know)

\[ \begin{array}{ccc}
    & 3d & 4p \\
    & 3p & 4s \\
    & 3s & \\
    2p & 2s & 1s
\end{array} \]

\[ \begin{array}{ccc}
    & 3d & 4p \\
    & 3p & 4s \\
    & 3s & \\
    2p & 2s & 1s
\end{array} \]

\[ \begin{array}{ccc}
    & 3d & 4p \\
    & 3p & 4s \\
    & 3s & \\
    2p & 2s & 1s
\end{array} \]

After writing the electron configuration, circle the valence electrons in the electron configuration.

\[ \text{e}^- \text{ config. for V} = \]

\[ \text{e}^- \text{ config. for As} = \]

\[ \text{e}^- \text{ config. for Cr} = \]

\[ ^{29}_{\text{Cu}} \) (also an exception, how?) \[ ^{16}_{\text{S}^{2-}} \]

\[ \begin{array}{ccc}
    & 3d & 4p \\
    & 3p & 4s \\
    & 3s & \\
    2p & 2s & 1s
\end{array} \]

\[ \begin{array}{ccc}
    & 3d & 4p \\
    & 3p & 4s \\
    & 3s & \\
    2p & 2s & 1s
\end{array} \]

\[ \begin{array}{ccc}
    & 3d & 4p \\
    & 3p & 4s \\
    & 3s & \\
    2p & 2s & 1s
\end{array} \]

After writing the electron configuration, circle the valence electrons in the electron configuration.

\[ \text{e}^- \text{ config. for Cu} = \]

\[ \text{e}^- \text{ config. for S}^{2-} = \]

\[ \text{e}^- \text{ config. for Fe}^{3+} = \]
How do the quantum numbers fit in?

Principal quantum number (n) \rightarrow “which apartment building?” \rightarrow n = 1, 2, 3… ∞
Angular quantum number (\ell) \rightarrow “which floor?” \rightarrow \ell = 0, 1, 2, 3…(n-1)
Magnetic quantum number (m_\ell) \rightarrow “which room?” \rightarrow m_\ell = +\ell …+2, +1, 0, -1, -2…+\ell
Spin quantum number (m_s) \rightarrow “which roommate?” \rightarrow m_s = + \frac{1}{2}, - \frac{1}{2}

1. The E diagram on the left is for the ground state of which element?
2. Give the quantum numbers of these 6 electrons.
3. Give the quantum numbers for these orbitals.
4. Give the quantum numbers for each of these 2 electrons.
5. What are the possible values of m_\ell for electrons with n = 3, \ell = 1?
6. Where would you find the electron with this set of quantum numbers: \ n = 3, \ell = 2, m_\ell = 0 and m_s = + \frac{1}{2}
7. What is wrong with this set of quantum numbers?
   n = 2, \ell = 2, m_\ell = 0 and m_s = + \frac{1}{2}

8. How many electrons can have n = 3, \ell = 2 and m_\ell = 1?
9. What is the maximum number of electrons that can have n = 4, \ell = 3?
10. What is the maximum number of electrons that can have n =3?
11. What is the number of electrons that ground state phosphorus has with n = 3, \ell =1?
    n = 3, \ell =2?
12. For each of the following sets of quantum numbers, state whether it is allowed:
   a) n = 4, \ell = 4, m_\ell = 0, m_s = - \frac{1}{2}
   b) n = 0, \ell = 0, m_\ell = 0, m_s = + \frac{1}{2}
   c) n = 3, \ell = 2, m_\ell = -2, m_s = + \frac{1}{2}
   d) n = 2, \ell = 0, m_\ell = +1, m_s = - \frac{1}{2}
Obtaining Electronic Configurations from the Periodic Table

<table>
<thead>
<tr>
<th>1</th>
<th>H / He</th>
<th>2s</th>
<th>3s</th>
<th>4s</th>
<th>5s</th>
<th>6s</th>
<th>7s</th>
<th>He</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>2s</td>
<td>3s</td>
<td>4s</td>
<td>5s</td>
<td>6s</td>
<td>7s</td>
<td>2p</td>
<td>3p</td>
</tr>
</tbody>
</table>

Lanthanide Series
Actinide Series

The diagram above will **not** be provided on an exam. A regular periodic table will be provided as usual.

Commit to memory:
1. Locations of s, p, d, f
2. Numbering system of each section.

Each square in the table above stands for an electron in an orbital.
Each element has all the electrons of the previous element, plus one more.

Give the electron configuration of each element below and **circle** the outershell electron(s). You may use **only** the periodic table.

<table>
<thead>
<tr>
<th># of valence e⁻</th>
<th>Group #</th>
</tr>
</thead>
<tbody>
<tr>
<td>6  C</td>
<td>1s² 2s² 2p²</td>
</tr>
<tr>
<td>7  N</td>
<td>1s² 2s² 2p³</td>
</tr>
<tr>
<td>8  O</td>
<td>1s² 2s² 2p⁴</td>
</tr>
<tr>
<td>9  F</td>
<td>1s² 2s² 2p⁵</td>
</tr>
<tr>
<td>10 Ne</td>
<td>1s² 2s² 2p⁶</td>
</tr>
<tr>
<td>17 Cl</td>
<td>____________</td>
</tr>
<tr>
<td>23 V</td>
<td>____________</td>
</tr>
<tr>
<td>34 Se</td>
<td>____________</td>
</tr>
<tr>
<td>24 Cr</td>
<td>____________</td>
</tr>
<tr>
<td>29 Cu</td>
<td>____________</td>
</tr>
<tr>
<td>Fe³⁺</td>
<td>____________</td>
</tr>
</tbody>
</table>

You should be able to write electron configurations for all the elements (atoms and ions, and exceptions such as Cu and Cr).
1. Using the core method, give the e\(^{-}\) configuration of…

\[ \text{Sn} \quad \text{Ba} \quad \text{P} \quad \text{Fe}^{3+} \]

2. Fill in the electrons. Give the full e\(^{-}\) config. and circle the valence e\(^{-}\).

\[
\begin{array}{cccc}
\text{Na} & 3d & 4s & 3p \\
 & 2p & 3s & 1s \\
\end{array}
\begin{array}{cccc}
\text{Cl} & 3d & 4s & 3p \\
 & 2p & 3s & 1s \\
\end{array}
\]

\[
\begin{array}{c}
\text{Na}^+ \\
\text{Cl}^- \\
\end{array}
\]

3. Give the e\(^{-}\) configuration of only the valence e\(^{-}\):

\[
\begin{array}{c}
\text{S} \\
\text{S}^2- \\
\text{K} \\
\text{K}^+ \\
\end{array}
\begin{array}{c}
\text{which resembles } \text{____} \\
\text{which resembles } \text{____} \\
\end{array}
\]

\[
\begin{array}{c}
\text{S}\text{, K}^+, \text{Ar} \text{ are said to be isoelectronic.}
\end{array}
\]

4. Name two other ions that are isoelectronic with them: ______ and ______.

5. a) Are Br\(^-\) and Cl\(^-\) isoelectronic? Yes or No

   b) Are Ba\(^{2+}\) and Cl\(^-\) isoelectronic? Yes or No

An atom is said to be “paramagnetic” if it contains unpaired electrons.
It is said to be “diamagnetic” if it does not contain any unpaired electrons.

Using ONLY a periodic table, answer the following questions:

6. Is silicon paramagnetic or diamagnetic?

7. How many unpaired electrons does sulfur have? Ans. ____ Is it paramagnetic or diamagnetic? Circle one.

8. How many unpaired electrons does the chromium atom have? Ans. ____

9. How many unpaired electrons does Fe\(^{3+}\) have? Ans. ____