EXPERIMENT 21.

Part 1. Bohr’s Planetary Model

Name: ____________________________
Section __________

The “Planetary” model of the atom has electrons “orbiting” around the nucleus in specific “orbits” or Energy levels. If an electron undergoes a transition from a high to a lower level, light is emitted with energy equal to the difference in the energies of the two energy levels. The transition is represented as an arrow. The energy of the emitted light is inversely proportional to its wavelength ($\lambda$) or directly proportional to its frequency ($\nu$).

\[ E = h \nu = h \frac{c}{\lambda} \]

$\lambda \times \nu = c$ \hspace{1cm} ($c = 3.00 \times 10^8$ m/sec) \hspace{1cm} ($h$ is Planck’s constant $= 6.63 \times 10^{-34}$ J sec, $\nu$ has the unit sec$^{-1}$, also called “Hertz” or Hz, while $\lambda$ in this equation is given in meters. $\lambda$ by itself is usually expressed in nm ($10^{-9}m$) or Å (Ångström $= 10^{-10}m$).

Exercises: use extra paper as necessary

1. Orange light emitted from glowing mercury vapor has a frequency of about $4.80 \times 10^{14}$ sec$^{-1}$. Calculate the wavelength of this light in m, nm, and Å.

2. The energy of light (and any other electromagnetic energy) is given by Planck’s equation: $E = h \nu$ or $h \frac{c}{\lambda}$. Calculate the energy in Joules, of the yellow light in problem 1.

3. Light can be characterized as a collection of particles called “photons”. The calculation in problem 2 gives the energy of one photon of yellow light. Calculate the energy of a mole of photons (use Avogadro’s #), and express the answer in kJ / mole.
In the lab: Your instructor will demonstrate, or you will follow a simple procedure to illustrate how light is emitted from atoms with electrons in higher energy levels.

a) Simple flame tests involve imparting energy to atoms, and allowing the energy to come back out as light of various colors. Describe the various colors of light emitted from solutions of the following ions when exposed to a flame.

Na$^+$ ___________________ K$^+$ ___________________ Ca$^{2+}$ ______________

Li$^+$ ___________________ Sr$^{2+}$ ___________________ Ba$^{2+}$ ______________

b) Examine light emitted from helium and hydrogen in spectrosopes. A spectroscope takes light from a glowing gas and separates it into specific colors with characteristic wavelengths.

When white light is passed through a prism or diffraction grating, one can separate violet light ($\lambda = 400$-420 nm), blue (~ 440-460 nm), green (460-500 nm), yellow (500 – 600 nm), orange (600 – 650 nm) and red (up to 700 nm) light.

The prism or diffraction grating in the spectroscope projects the specific colors onto a scale. The scale in the spectrosopes is given in cm x 10$^{-5}$. Since 1 cm = 10$^7$ nm, a scale reading on our spectrosopes of 1 is equivalent to 100 nm; a scale reading of 4 is equivalent to 400 nm. Record the wavelengths and colors of the lines coming from the glowing hydrogen gas and the glowing helium gas. (Convert the scale to nm)

<table>
<thead>
<tr>
<th>color</th>
<th>$\lambda$</th>
<th>color</th>
<th>$\lambda$</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>___________ nm</td>
<td>He</td>
<td>___________ nm</td>
</tr>
<tr>
<td></td>
<td>___________</td>
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<td>___________</td>
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</table>

Note – you may not see all 4 lines - the one close to 400 nm is dim

Niels Bohr was able to calculate the wavelengths of the light emitted by glowing hydrogen gas, for transitions between higher energy levels (n = 3, 4, etc) and the second energy level (n = 2). The formula is:

$$\lambda = \frac{91.2 \text{ nm}}{0.250 - \frac{1}{n^2}}$$

where n = 3 or 4 or 5 or 6

$\lambda$ (n = 3, 4, 5, 6), to obtain 4 wavelengths.

Compare these with the wavelengths of the light you saw in the spectroscope.

<table>
<thead>
<tr>
<th>calculated values</th>
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<tbody>
<tr>
<td>___________ nm</td>
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<td>___________ nm</td>
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<tr>
<td>___________ nm</td>
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<td>___________ nm</td>
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</tbody>
</table>
Part 2. The Quantum Mechanical Model

This model of the atom, has the electrons behaving as waves, and only certain waves are mathematically possible in a closed 3-D system such as the atom, which give rise to a description of probability of finding the electrons in certain regions of the atom. These probability distributions give rise to 3-D pictures called “orbitals”. The number and types of orbitals vary with the energy levels and sublevels (subshells) in the atom. Thus the principal energy levels (designated n = 1 or 2 etc) split into sublevels (or “subshells) of different energies, which further split into orbitals.

The first energy level remains unsplit – and contains a spherical-shaped “s” orbital. (1 s subshell has 1 orbital) The 2nd (n = 2) level is split into two subshells, the 2s (one orbital) and the 2p (three orbitals).

s orbitals are spherical

\[ s \]

the three p orbitals have dumbbell shapes and are oriented along an x,y,z axis

\[ p \]

The 3rd energy level has
A 3 s subshell (1 orb), a 3 p subshell (three orbs) and a 3 d subshell of five orbitals having cloverleaf shapes.

The 4th energy level is split into 4 subshells: 4s (one orb) 4p (three orbs) 4d (five orbs) and 4f (five orbs)
The 4f orbitals have complex shapes and there is no need to learn how to sketch them. The energies of all the orbitals are plotted, in what is called an “Aufbau” diagram:

<table>
<thead>
<tr>
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</thead>
<tbody>
<tr>
<td>7s</td>
<td>6d</td>
<td>5f ………actinides</td>
</tr>
<tr>
<td>6s</td>
<td>5d</td>
<td>4f …… lanthanides</td>
</tr>
<tr>
<td>5p</td>
<td>4d</td>
<td>54 …… [ Xe ]</td>
</tr>
<tr>
<td>5s</td>
<td>4s</td>
<td>#37 Rb = [ Kr ] 5s^1</td>
</tr>
<tr>
<td>4p</td>
<td>3d</td>
<td>#19 K = [ Ar ] 4s^1</td>
</tr>
<tr>
<td>4d</td>
<td>3s</td>
<td>#11 Na = [ Ne ] 3s^1</td>
</tr>
<tr>
<td>3p</td>
<td>2p</td>
<td>10 ….[ Ne ]</td>
</tr>
<tr>
<td>2p</td>
<td>2s</td>
<td>18 …[ Ar ]</td>
</tr>
<tr>
<td>3s</td>
<td>4s</td>
<td>36 …[ Kr ]</td>
</tr>
<tr>
<td>4p</td>
<td>6p</td>
<td>86 …[ Rn ]</td>
</tr>
<tr>
<td>5d</td>
<td>5f</td>
<td>one e’ goes into the 5d before the 4f subshell, then after the 4f is filled, the electrons fill the 5d subshell</td>
</tr>
<tr>
<td>7s</td>
<td>6d</td>
<td>Each noble gas configuration has a filled outer p subshell</td>
</tr>
</tbody>
</table>

Exercise  Predict the electron configuration of the element lead (#82), by filling with dots the energy levels (orbitals) in the above Aufbau diagram (max 2 per orbital as in the 1s)
2. Using dots (or arrows) to represent electrons. Fill in the following Aufbau diagrams.

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

germanium  manganese  bromide ion  vanadium (V) ion

Using subshell notation, \((1s^22s^2\text{ etc})\) write complete electron configurations

\[
\begin{align*}
\text{Ge} & \quad \_{\text{complete electron configuration}} \\
\text{Mn} & \quad \_{\text{complete electron configuration}} \\
\text{Br}^{1+} & \quad \_{\text{complete electron configuration}} \\
\text{V}^{+5} & \quad \_{\text{complete electron configuration}} \\
\end{align*}
\]

Now write abbreviated electron configurations using noble gas notation \([\quad]\) plus partially filled outer subshells. i.e. iron: \(1s^22s^22p^63s^23p^64s^23d^6\); abbreviated as \([\text{Ar}]4s^23d^6\)

\[
\begin{align*}
\text{Ge} & \quad \_{\text{abbreviated electron configuration}} \\
\text{Mn} & \quad \_{\text{abbreviated electron configuration}} \\
\text{Br}^{1+} & \quad \_{\text{abbreviated electron configuration}} \\
\text{V}^{+5} & \quad \_{\text{abbreviated electron configuration}} \\
\end{align*}
\]

3. Write the electron configurations of these elements and their ions:

\[
\begin{align*}
\text{Al} & \quad \_{\text{electron configuration}} \\
\text{Al}^{+3} & \quad \_{\text{electron configuration}} \\
\text{Cl} & \quad \_{\text{electron configuration}} \\
\text{Cl}^- & \quad \_{\text{electron configuration}} \\
\text{Ba} & \quad \_{\text{electron configuration}} \\
\text{Ba}^{+2} & \quad \_{\text{electron configuration}} \\
\text{As} & \quad \_{\text{electron configuration}} \\
\text{As}^{+3} & \quad \_{\text{electron configuration}} \\
\text{V} & \quad \_{\text{electron configuration}} \\
\text{V}^{+3} & \quad \_{\text{electron configuration}} \\
\text{Sb} & \quad \_{\text{electron configuration}} \\
\text{Sb}^{+3} & \quad \_{\text{electron configuration}} \\
\text{Sb}^{+5} & \quad \_{\text{electron configuration}} \\
\end{align*}
\]

4. Write symbols of three cations & three anions that are isoelectronic with neon: (isoelectronic means having the same number of electrons)

5. Write Lewis electron dot formulas showing the valence electrons of:

\[
\begin{align*}
\text{O} & \quad \_{\text{Lewis electron dot formula}} \\
\text{Se} & \quad \_{\text{Lewis electron dot formula}} \\
\text{P} & \quad \_{\text{Lewis electron dot formula}} \\
\text{Br} & \quad \_{\text{Lewis electron dot formula}} \\
\text{Ga} & \quad \_{\text{Lewis electron dot formula}} \\
\text{Sn} & \quad \_{\text{Lewis electron dot formula}} \\
[\text{Bi}]^{3-} & \quad \_{\text{Lewis electron dot formula}} \\
\text{K} & \quad \_{\text{Lewis electron dot formula}} \\
[\text{Ca}]^{+2} & \quad \_{\text{Lewis electron dot formula}} \\
\text{Cs} & \quad \_{\text{Lewis electron dot formula}} \\
\end{align*}
\]