**Honors Review Sheet – Quantum Mechanics Quiz**

Material that will be covered:

* How quantum mechanics differs from Bohr
* How spectroscopy works
* How the flame test worked
* How line spectra are formed
* The quantum numbers and what they represent
* Pauli exclusion principle, Hund’s rule, and aufbau principle
* Electron configurations
* Orbital filling diagrams

1. Describe three ways that the quantum mechanical model of the atom differs from the Bohr model.
   * **The Bohr model says that electrons are particles, while QM says they’re waves**
   * **Bohr model has one QN, QM has four**
   * **Bohr has circular orbitals around the nucleus, QM has irregularly shaped orbitals**
   * **Bohr says farther from the nucleus = more energy, QM doesn’t have the same direct correlation**
   * **Bohr says orbitals never overlap, QM says they do**
   * **Bohr says different orbitals can hold different numbers of electrons, QM says they can hold a maximum of two.**
2. Explain in detail the process by which an atom can be made to emit light and form a line spectrum.
   * **An electron is in a ground state orbital**
   * **Energy is added to the atom, which causes the electron to jump to an excited state**
   * **When the electron falls back down to the ground state, it gives off the energy it absorbed as a particular color of light.**
3. Why are four variables required in the quantum mechanical model of the atom, while only one is required for the Bohr model?
   * **The Bohr model of the atom could only predict the orbital energies/line spectra for hydrogen. The quantum model required additional variables to make it more widely applicable.**
4. What are the allowed values for the quantum numbers of an atom if n=4?
   * **l = 0, 1, 2, 3**
   * **ml = -3, -2, -1, 0, 1, 2, 3**
   * **ms = ½, -1/2**
5. What types of orbitals (out of s, p, d, f) can exist in the atom from problem 4?
   * **The l variable determines the types of orbitals that are allowed. Since l = 0 denotes an s-orbital, l=1 denotes a p-orbital, l = 2 denotes a d-orbital, and l=3 denotes and f-orbital, all four of these types of orbitals are allowed.**
6. What does the Pauli exclusion principle have to do with the spin quantum number?
   * **The Pauli exclusion principle says that no two electrons can have the same four quantum numbers in an atom. In order to ensure that the two electrons in an orbital have separate quantum numbers, the spin quantum number is used to distinguish between them.**
7. What is the relevance of Hund’s rule when writing orbital filling diagrams?
   * **Hund’s rule says that electrons prefer to be unpaired whenever possible. As a result, when writing an electron filling diagram, we always draw electrons as unpaired within orbitals before we start pairing them up.**
8. Write the electron configurations for Bi, U, and Pd.

*NOTE: In order to save myself a huge amount of time making superscripts when writing these energy diagrams, I have written them as a series of terms utilizing only normal writing. The correct version of each term has the number after the orbital type as a superscript. For example, the term I’ve written as “1s2” is actually written as 1s2.*

* + **Bi: 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6 5s2 4d10 5p6 6s2 4f14 5d10 6p3 (If you wish to write the abbreviated version, you’d start with [Xe] and then write only the terms from 6s2 onward.**
  + **U: 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6 5s2 4d10 5p6 6s2 4f14 5d10 6p6 7s2 5d4 (If you wish to write the abbreviated version, start with [Rn] and then write only the terms from 7s2 onward.**
  + **Pd: 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6 5s2 4d8 (If you wish to write the abbreviated version, start with [Kr] and write only the terms from 5s onward.**

1. Draw the orbital filling diagram for Pd:



![Diagram

Description automatically generated]()