Unit 5: Atomic Models and Periodic Table

Study Guide ***KEY***

Historic Atomic Models:

1. Describe and diagram each of the following historic atomic models: Dalton, Thompson, Rutherford, and Bohr.
   Dalton: This model should show only spherical atoms with no subatomic particles. Two atoms of different types could be shown as well, to represent that different elements have different atoms.

   Thompson: This model is spherical with electrons randomly placed throughout. There should also be a positive charge represented in some way, to balance out the negative electrons, although Thompson didn't know what the positive charge was.

   Rutherford: This model should have a positive nucleus containing more than 99% of the mass of the atom. The electrons are randomly dispersed around the nucleus in what is mostly "empty space".

   Bohr: This model has protons and neutrons in the nucleus, with electrons travelling in distinct paths called orbits.

2. What experiment did JJ Thompson do and how did it change Dalton's ideas about atoms?
   Thompson experimented with a cathode ray tube. He passed electricity through this glass tube, and exposed the beam to a magnet. The negative end of the magnet repelled the beam, showing that the electricity was made of negatively charged particles. He had discovered electrons, forcing Dalton's atom to incorporate electrons.

3. What experiment did Rutherford do and how did this change Thompson's ideas about atoms?
   Rutherford performed the Gold Foil Experiment. He shot alpha particles at a piece of gold foil, observing how they interacted with it. A large percentage of the alpha particles went straight through the atoms in the foil, showing that an atoms must be made of mostly empty space. About 1 % were deflected backwards, showing that there must be some part of the atom massive enough to cause the deflection. He had discovered the nucleus, and forced Thompson's model to incorporate a nucleus.

4. How did the Bohr model improve upon Rutherford's model?
   The Bohr model introduced electron organization. Bohr put electrons in distinct, rotating paths called orbits, rather than the random placement they still had in Rutherford's model.

5. Even though Bohr's model wasn't completely correct, it was fairly accurate, and we still use it for some purposes to this day. Critique Bohr's model in terms of how he was correct, and how he was incorrect.
   Bohr's model was relatively correct in terms of the nucleus and the idea of electrons having energy levels. It was incorrect in saying that electrons could "orbit" the nucleus in a distinct path, or ring. Electrons do have energy levels that they belong to, but within those energy levels are regions of space where the electron is most likely to appear. These are called orbitals.

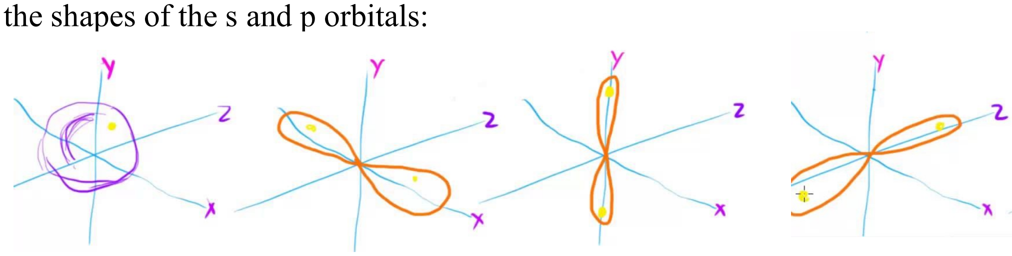
Quantum Mechanical Model:
What is another name for this model?
Electron cloud model

What did Schrodinger call the area of space where electrons are **most likely** to be found?
orbitals

List all the orbital types that would be found in energy level 4.
s, p, d and f orbitals

What is the maximum number of electrons that can fit in the 3rd energy level?
The 3rd energy level has s, p and d orbitals. s orbitals hold 2 electrons, p orbitals can hold 6 electrons and d orbitals can hold 10 electrons, so 18 total electrons (2 + 6 + 10) can fit in the third energy level.

Draw the shapes of the s and p orbitals:



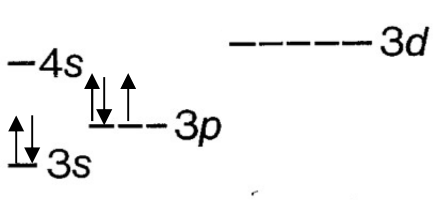
What is the maximum number of d orbitals allowed in any one principal energy level?
d orbitals hold 10 electrons, meaning that there must be 5 separate d orbitals.

How many sublevels are in the third principal energy level?
The term sublevel is refers to the types of orbitals in a particular energy level. The third energy level has s, p, and d orbitals, so 3 different sublevels.

Electron Configurations:
The Aufbau principle tells us that electrons enter **what** orbital first?
Lowest energy orbitals first

The Pauli exclusion principle tells us that an orbital may hold how many electrons?
2 electrons at most

Hund's rule tells us to do what with electron configurations?
When electrons are filling multiple orbitals of equal energy, they spread out over all the open orbitals, before pairing up.



What rules does the following configuration violate?
How would you correct it?
This violates hund's rule. The first orbital in the 3p's would not have two electrons because the third orbital is still open. Each orbital would only have 1 electron.

Write the unabbreviated electron configuration for the following:
Oxygen: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
Americanium: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2\ 5f^7$

Write the abbreviated electron configuration for the following:
Magnesium: ***[Ne] $3s^2$***
Platinum: ***[Xe] $6s^2$ $4f^{14}$ $5d^8$***

Are the following configurations correct or in-correct?
$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$ ***incorrect – 4s comes after 3p, not 3d***
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9 4p^5$ ***incorrect – the 3d orbitals hold 10 electrons, not 9***
[Pb] $7s^2$ ***incorrect – only noble gases are used in the brackets***

Periodic Table:
1. Who are the two main developers of the periodic table and how did they arrange it? Whose model do we use today? ***Mendeleev arranged atoms by atomic weight, Moseley arranged atoms by atomic #***
2. What group number are the noble gases in? ***group 8 or group 8A***
3. What group is Potassium in? ***alkali metals*** What type of orbital do the electrons fill to? ***s^1***
4. What group is Osmium (Os) in? ***transition metals*** What type or orbital do the electrons fill to? ***d***
5. What types of orbitals do the representative elements fill to? ***s and p orbitals***
6. The transition metals fill to what type of orbital? ***d orbitals***
7. List the atoms in order of increasing size. Pb, Sn, C, Si, Ge ***C, Si, Ge, Sn, Pb***
8. List the atoms in order of decreasing ionization energy. Ba, Lu, Hg, W, Bi **Bi, Hg, W, Lu, Ba**
9. List the atoms in order of increasing electronegativity. Xe, Rn, He, Ar, Kr ***Rn, Xe, Kr, Ar, He***
10. What is the halogen with the largest size? ***astatine***
11. What is the most reactive element in Oxygen's column? ***oxygen***

12. Fully explain why Aluminum has a larger atomic radius than Chlorine. Use complete sentences and diagrams are welcome.
Chlorine is smaller than aluminum because it has more protons in it's nucleus. This means that chlorines electrons are more attracted to the nucleus, and therefore pulled closer to it, making the radius smaller.

13. Fully explain why Barium has a larger atomic radius than Magnesium. Use complete sentences and diagrams are welcome.
Barium has a larger atomic radius than magnesium because it fills to the 6th energy level, compared to magnesium which only fills to the 3rd energy level.

13. Fully explain Fluorine has the highest ionization energy of all the halogens. Use complete sentences and diagrams are welcome.
Fluorine has the highest ionization energy because it is the smallest halogen. Smaller atoms have electrons that are closer to the nucleus. They are therefore more attracted to the nucleus, requiring higher energies to be removed.

14. Explain how the flame tests provide evidence supporting the existence of energy levels for electrons. The following words should be used in your response: a) ground state atom, b) excited state atom, c) energy, d) electron jumps, e) electron drops, f) emitted light

The process begins with an atom in the ground state, meaning it's electrons are in their "normal", low energy orbitals. When energy, usually in the form of heat, is added to the atoms, the electrons "jump" to higher energy levels. The atom is now in an "excited" state. The electrons do not remain at these higher levels permanently. They quickly drop back to the orbitals they came from, emitting light in the process. We observe this emitted light in the form of different flame colors.