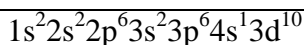


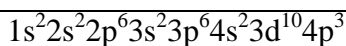
Lecture 3
Atomic Theory III
Tutorial

- 1) What is the electron configuration for copper in its ground state? (long form)

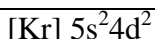


The top row of the d-block is in the $n = 3$ shell.

- 2) What is the ground state electron configuration of arsenic? (long form)

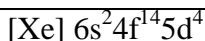


- 3) What is the ground state electron configuration of zirconium? (short form)



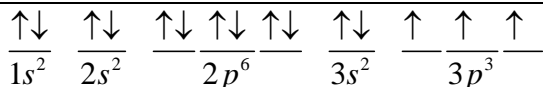
Put the noble gas from the period above the element in question in square brackets. This notation stands for the ground state electron configuration of that noble gas. Starting from the left side of the periodic table, in the period that contains the element in question, add the additional subshells (and the number of electrons in those subshells) that exist in the ground state of that element.

- 4) What is the ground state electron configuration of tungsten? (short form)



The top row of the f-block is in the $n = 4$ shell. The f-block fits in between the s-block and the d-block. The top row of the f-block starts after 6s. The third row of the d-block is in the $n = 5$ shell.

- 5) Draw an orbital diagram for phosphorus in its ground state.



According to Hund's Rule, one electron is added to each orbital in a subshell before a second can be added. This is why we end up with one electron in each of the orbitals of the 3p subshell in phosphorus. If two electrons were added to one orbital and one was added to another, as is indicated here: $\frac{\uparrow\downarrow \uparrow}{3p^3}$, it would be incorrect. This is due to

the fact that $\frac{\uparrow \uparrow \uparrow}{3p^3}$ is at a lower energy level. Nature always favors systems that are at lower energy levels.

d. I or I⁻

I has the smaller radius. Both I and I⁻ have the same number of protons, but I has one less electron. This causes the protons to pull on the electrons in I with greater forces of attraction. The greater forces of attraction caused by the larger effective nuclear charge causes the radius of I to be smaller.

8) Which of the following are isoelectronic:

a. Ne and O²⁻

These are isoelectronic. O²⁻ gained two electrons, giving it a total of ten electrons. Ne also has ten electrons. Both of these species have the same ground state electron configurations, as both have the same number of electrons.

b. K⁺ and Cl

These are NOT isoelectronic. K⁺ lost one electron, giving it a total of eighteen electrons. Cl has seventeen electrons. These species do not have the same ground state electron configurations, as they have different numbers of electrons.

9) Rb⁺ and Br⁻ are isoelectronic.

a. Which ion has the smaller radius?

Rb⁺

b. Explain why the radii of these two ions are different sizes.

Both species have the same ground state electron configurations, and thus, the same number of electrons. Rb⁺ has 37 protons, while Br⁻ only has 35 protons. Because Rb⁺ has more protons, its nucleus applies a greater force of attraction on its electrons. This greater force of attraction pulls the electrons closer to the nucleus, thereby giving Rb⁺ a smaller radius.

10) Explain each of the following occurrences by referencing the structure of the atoms in question (energy levels, orbitals, protons, etc.).

a. The first ionization energy of Beryllium is greater than the first ionization energy of Boron.

This is an exception to the rule. Normally ionization energy increases as we move from left to right across a period in the periodic table, because the effective nuclear charge increases. The reason for this exception has to do with the electron configurations of these elements. Be has an electron configuration of 1s²2s², and B has an electron configuration of 1s²2s²2p¹. It requires less energy to pull the only electron from the p-orbital in Boron, which is a higher energy subshell; than it does to pull an electron from the full s-orbital in Beryllium, which is a lower energy subshell. The lone electron in the p-orbital of Boron is at a much higher energy than those in the full s-orbital of Beryllium.

- b. The first ionization energy of Li is greater than the first ionization energy of K.

K has a larger atomic radius than Li, due to the fact that the valence electrons of K are in a 4s orbital and those of Li are in a 2s orbital. As the radius of K is larger, its valence electrons are held with a smaller force of attraction, and thus, it requires less energy to remove them.

- c. The first ionization energy of Ca is less than the first ionization energy of Se.

Ca has a larger atomic radius than Se. The valence electrons of both elements are in the same shell ($n = 4$). However, Se has a smaller radius. It has more protons in its nucleus that provide a greater force of attraction on its electrons. Because the force of attraction on the electrons in Ca is less than that of Se, it requires less energy to remove an electron from Ca.

- d. Which structure has the highest fourth ionization energy: Carbon or Boron? Justify your answer.

Boron has the highest fourth ionization energy. The fourth electron that is removed from Boron is taken from the $n = 1$ shell, whereas the fourth electron that is removed from Carbon is taken from the $n = 2$ shell. The distance between the nucleus and the $n = 1$ shell electrons is much less than the distance between the nucleus and the $n = 2$ electrons. The reduced distance means that the attractive forces between the protons and the $n = 1$ electrons of Boron are greater. The greater the attractive force, the more energy that is required to remove an electron.