

Fall 2008 CH301 Worksheet 3 Answer Key—Last half of Chapter 1 in preparation for Quiz 2

- List the following orbital in order of increasing energy in the hydrogen atom: 1s, 10s, 4p, 4s, 3d
The correct order would be: $1s < 4s < 3d < 4p < 10s$. The first four can be read right off the periodic table, and 10s has a much higher principle energy level ($n = 10$) and is thus highest in energy.
- What is the ground state electron configuration of S, Sulfur in both the long and short hand notation?
The long hand notation would be: $1s^2 2s^2 2p^6 3s^2 3p^4$. The last noble gas before Sulfur is Neon, whose electron configuration is $1s^2 2s^2 2p^6$, therefore the short hand notation would be $[\text{Ne}]3s^2 3p^4$.
- What is the ground state electron configuration of metallic Zinc, Zn, written in short hand notation?
The last noble gas before Zinc in Argon, so $[\text{Ar}]4s^2 3d^{10}$
- What is the ground state electron configuration of bromide ion, Br^- , written in long hand notation?
 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
- Which of the following pairs of species have the same electron configuration?
 - Na^- and Ne (12 and 10 electrons respectively)
 - Mg^{2+} and Ar (10 and 18 electrons respectively)
 - Cl^- and Li^+ (10 and 18 electrons respectively)
 - Ca^+ and Ar (19 and 18 electrons respectively)
 - K^+ and Ar (18 and 18 electrons respectively)**The only thing that determine electron configuration is the number of electrons.
- How many unpaired electrons are in the ground state of Cu, copper? For an extra challenge, which subshell contains the unpaired electron.
There is just one unpaired electron in the ground state of copper. The unpaired electron in Cu is in the 4s orbital. The electron configuration for Cu is $[\text{Ar}]4s^1 3d^{10}$. It is an exception.
- How many electrons are there in the electron shell with principle quantum number $n = 3$?
There are 18. Principle energy level 3 contains a 3s, 3p and 3d subshell, holding 2, 6 and 10 electrons respectively, for a total of 18. In general number of electrons in a shell is equal to $2n^2$.
- Explain why an electron in a shell farther from the nucleus "feels" less attraction to the nucleus than one closer to it. How does this explain why effective nuclear charge decreases down the periodic table but increases going to the right?
The inner shell electrons "shield" the valence electrons from the full Coulombic attraction of the nucleus. Going down a column or family in the periodic table, each element's valence electrons experience the same effective nuclear charge (ENC, aka Z_{eff}) because the difference in the number of protons and the number of inner shell electrons is constant ($Z_{\text{eff}} = Z - S$). Going across a row from left to right, however, the positive charge of the nucleus increases while the number of inner shell electrons remains the same, so valence electrons experience a greater effective nuclear charge.
- How does the effective nuclear charge of an atom influence its size? Use this explanation to explain the periodic trends in atomic radii - do the atoms get larger at the top or bottom? The left or right?

Going down columns or families in the periodic table, the effective nuclear charge (Z_{eff} , aka ENC) is constant, but because the principle energy level increases, the electrons must be farther away from the nucleus and thus atomic radius increases. Going across a period from left to right, we have greater ENC, in other words more attractive force on the outermost electrons, and so atomic radius decreases.

10. Applying knowledge from question 2, which of the following pairs of atoms is larger?

A. C or Ge

Ge is beneath C in the same group, therefore it is larger.

B. Sb or Te

Sb is to the left of Te in the same period, therefore it is larger.

C. Fr or Na

Fr is beneath Na in the same group, therefore it is larger.

D. N or F

N is to the left of Te in the same period, therefore it is larger.

11. What is ionization energy? Explain how ENC effect ionization energy? What are the periodic trends for ionization energy?

Ionization energy (IE) is the energy needed to remove a valence electron from an atom. When ENC is higher, the electrons are more tightly held so it requires more energy to remove them. So going across a period, IE increases just as ENC does. Going down a group, ENC is a constant, but remember the valence electrons are also further away (less tightly held) and thus IE decreases. IE increases across and decreases down.

12. Define electron affinity. Would you expect its periodic trends to match those of ionization energy? Why or why not? What are the periodic trends for electron affinity?

Electron affinity (EA) is the energy released when adding an electron to a neutral atom. This EA trend measures the opposite process of IE, but because we're measuring the energy released rather than the energy put in, the trend is actually identical – it increases across and decreases down. This makes sense. ENC increases across, so F for instance, really wants an electron.

13. Define electronegativity. Explain its periodic trends. What would you expect to be the most electronegative atom?

Electronegativity is the ability of an atom to attract electrons to itself. So going across a period, electronegativity increases just as ENC does. Going down a group, ENC is a constant, but remember principle energy level increases and so electronegativity decreases – the atom has trouble holding on to those far away electrons to begin with, so attracting an additional electron is difficult.

14. Write the electron configurations for V, Cr, and Mn.

V [Ar]4s²3d³

Cr [Ar]4s¹3d⁵ (Cr uses a 4s electron to make its 3d subshell half filled)

Mn [Ar]4s²3d⁵

15. What is the correct electron configuration for Zn^{+} ?

$[\text{Ar}]4s^13d^{10}$. This is because d subshells can borrow electrons from s subshells to be filled (borrows one or both s electrons) or half filled (borrows one s electron but never both).

16. How many electrons does Tl^{3+} have in its 6s subshell?

It has zero electrons in its 6s subshell. Both have been borrowed to leave the 5d subshell filled. In other words the electron configuration of Tl^{3+} is $[\text{Xe}]6s^05d^{10}$.

17. Write the electron configurations for Ni, Cu, and Zn.

Ni $[\text{Ar}]4s^23d^8$ (Ni uses both 4s electrons to make its 3d subshell filled)

Cu $[\text{Ar}]4s^13d^{10}$ (Cu uses a 4s electron to make its 3d subshell filled)

Zn $[\text{Ar}]4s^23d^{10}$

18. Which of the 3 rules for electron configuration do the *d*-block exceptions modify?

The Aufbau principle is modified by these exceptions. Hund's rule and the Pauli exclusion principle are left intact.

19. Which of the following statements are true?

A. A s^1d^5 configuration is at a higher energy than a s^2d^4 configuration.

B. A s^1d^5 configuration is more stable than a s^2d^4 configuration.

C. All ions have the same electron configuration as the neutral element with the same number of electrons.

D. The apparent exceptions to Aufbau can be understood through the tendency to form filled and half-filled subshells over other configurations.

E. Hund's rule predicts the s^1d^5 configuration instead of s^2d^4 , because it avoids pairing up two electrons in the s subshell.

Statements A and B are mutually exclusive, and because s^1d^5 configurations are what are observed experimentally, it is safe to say that they are more stable (lower energy) than s^2d^4 configurations. Statement C is true – the only thing that determines electron configuration is the number of electrons. Statement D is true – filled and half filled subshells are more stable than their partially filled counterparts. Statement E is false – Hund's rule never suggests we avoid pairing electrons, just that we wait to do it until a subshell is at least half full.

20. What is *always* true about ground-state electron configurations in regards to energy?

The electron configurations we actually observe are the configurations with the lowest possible energy.