

CH 221 Chapter Six Part II Concept Guide

1. Writing Electron Configurations

Question

What is the complete electron configuration of the zirconium atom?

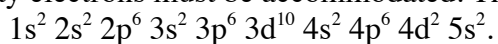
Approach

With increasing atomic number, electrons occupy the subshells available in each main energy level in order, with few exceptions:

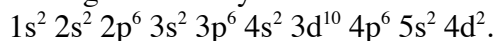


Solution:

Forty electrons must be accommodated. The total electron configuration for Zr is:



This configuration may also be written as:



2. Writing Electron Configurations

Question

What is the complete electron configuration for the arsenic atom?

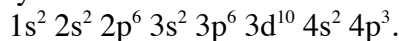
Approach

With increasing atomic number, electrons occupy the subshells available in each main energy level in order, with few exceptions:

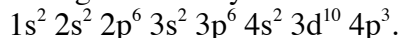


Solution:

Thirty-three electrons must be accommodated. The total electron configuration for As is:



This configuration may also be written as:



3. Writing Electron Configurations

Question

What is the noble gas notation for the electron configuration for the rubidium atom?

Approach

With increasing atomic number, electrons occupy the subshells available in each main energy level in specific order, with few exceptions:



The noble gas notation substitutes the symbol of the noble gas for the corresponding noble gas core in the electron configuration.

Solution:

The noble gas just prior to Rb in the periodic table is Kr. The noble gas notation for the electron configuration for Rb is: $[\text{Kr}]5s^1$.

4. Writing Electron Configurations

Question

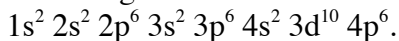
What is the complete electron configuration for the Br^- ? With which element in the periodic table is it isoelectric?

Approach

With increasing atomic number, electrons occupy the subshells available in each main energy level in specific order:

**Solution:**

The anion Br^- is formed by the addition of 1 electron to the lowest energy orbital that has a vacancy. The electron configuration for Br^- is:



Cl^- is isoelectric with krypton.

5. Electron Configurations of Transition Metal Cations

We will look at iron, a transition metal, and two of its cations, and compare their electron configurations. We will compare what we would expect from main group behavior against the real behavior of the cations. We'll examine the real behavior closely to understand why it occurs, and then come to some conclusions about why the real behavior differs from our expectations.

We will be examining Fe, Fe^{2+} , and Fe^{3+} . The electron configuration of neutral Fe is $[\text{Ar}]4s^23d^6$.

Question: What will be the electron configurations of Fe^{2+} and of Fe^{3+} ?

Expectation (based on main group behavior): If Fe loses electrons to form cations, we might expect the electrons removed to be those that were added last. We would expect:



Real Behavior:



Elaboration: Most transition metal atoms have electron configurations with two electrons in an s orbital, and some electrons in the d subshell of the shell below (with a lower value of n) that of the s orbital.

When transition metal atoms lose electrons to form ions, the first electrons lost are those in the s orbital. The electron configuration of Fe^{2+} is $[\text{Ar}]3d^6$. The 4s electrons have been lost, the 3d electrons remain. If a third electron is lost, forming Fe^{3+} , it will be removed from the 3d subshell, forming an ion with an electron configuration $[\text{Ar}]3d^5$.

Explanation: In an atom of Ca, the 4s subshell fills in preference to the 3d subshell and the atom is lower in energy with an $[\text{Ar}]4s^2$ electron configuration. Having a filled s subshell is favored for transition metal atoms as well, but this does not hold true for transition metal cations. In these cases, it is energetically favorable to have the d orbital subshell fill in preference to the s subshell. This implies that orbital subshells shift both in absolute energy as well as in energy relative to one another when ions form. This is the case.

Additional problems:

- 1) Give the electron configurations for O, S, and Se in spectroscopic notation. In what way are they similar?
- 2) Give the noble gas notation electron configuration for samarium, Sm, which is in the lanthanide series. What subshell remains only partially filled? Is a samarium atom expected to be diamagnetic or paramagnetic?
- 3) Give three atoms or ions that have an electron configuration of: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$.

1. Answer:

O: $1s^2 2s^2 2p^4$

S: $1s^2 2s^2 2p^6 3s^2 3p^4$

Se: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$

They are similar in that they each have an outermost p subshell that contains four electrons. Each of the elements is located in group 4A of the periodic table. In general, elements that are found in the same periodic group have similar outermost electron configurations.

2. Answer:

$[\text{Xe}]6s^2 4f^6$

The 4f subshell is partially filled, as is expected for an f-block element. The 4f subshell has 6 electrons (the overall subshell, with seven orbitals, can hold 14 electrons). Sm has atoms among the most strongly paramagnetic of all the elements because its atoms have 6 unpaired electrons.

3. **Answer:** Notice that this electron configuration has an outermost 5p subshell that is filled. It therefore represents the noble gas element in period 5 of the periodic table. This is Xe.

Atoms that can form anions with the same electron configuration will be found to the left of Xe in the periodic table. Therefore, I and Te^{2-} will each have this electron configuration. Atoms that can form cations with the same electron configuration will be the next two elements on the periodic table.

Therefore, Cs^+ and Ba^{2+} will each have this electron configuration. These five species, Te^{2-} , I, Xe, Cs^+ , and Ba^{2+} are all isoelectronic. They each have 54 electrons but, due to their differing number of protons, they have different charges.

6. Atomic Radii

Problem

Consulting the periodic table, decide whether the first atom in each of the following pairs is larger, smaller, or similar in atomic radius to the second atom:

- (a) Si, Pb (b) Cs, Pb (c) Rh, Ir (d) Ti, V.

Approach

Atomic size increases moving down a group, and decreases moving left to right along a period.

Solution:

- (a) Silicon is the second element in the carbon family, and lead is the fifth element. Silicon atoms should be smaller as size increases down a representative periodic group.
- (b) Cesium and lead are in the same period, with lead further to the right in the period. Size decreases across the periods, and therefore cesium atoms should be larger than lead atoms.
- (c) Rhodium and iridium are the second and third members of a d-transition metal group. Elements in the second and third transition series are very similar in size due to the lanthanide contraction, therefore, these atoms should be similar in size.
- (d) Titanium and vanadium are adjacent elements in the same d-transition period. The decrease in size across the periods for transition elements is gradual. Atoms of these two elements should be similar in atomic radius.

7. Ionic Radii

Problem

Place the following species in order of increasing radius:



Approach

For species having the same number of protons, the more electrons, the larger the species.

Solution:

The positively charged chlorine atom is smaller than neutral chlorine; the former has one fewer electron than neutral chlorine. The negatively charged chlorine atom is the largest of all three; it has one additional electron than neutral chlorine.



8. Ionic Radii

Problem

Arrange the following ions in order of increasing radius:



Approach

Ionic size increases moving down a group, and decreases moving left to right along a period. For isoelectric species, the greater the number of protons, the smaller the species.

Solution:

Mg^{2+} , Na^+ , F^- all have the same number of electrons, thus they are similar in size. Mg^{2+} , however, has the largest number of protons, therefore it is the smallest of these three ions. Likewise, S^{2-} is larger than Cl^- because it has fewer protons. S^{2-} and Cl^- have an additional shell relative to Mg^{2+} , Na^+ , and F^- , therefore S^{2-} and Cl^- are both larger than these three ions. Be^{2+} has one fewer shell, and is therefore smaller than Mg^{2+} , Na^+ , and F^- .

