Learning Objectives

- Determine the electron configuration of ions
- Justify the observed charge of ions to their electronic configuration
- Define paramagnetism and diamagnetism
- Justify the anomalies of the electron configurations in transition metals using magnetism experimental data

Electronic Configurations of Cations and Anions

The way we designate electronic configurations for cations and anions is essentially similar to that for neutral atoms in their ground state. That is, we follow the three important rules: Aufbau’s Principle, Pauli-exclusion principle, and Hund’s Rule. The electronic configuration of cations is assigned by removing electrons first in the outermost p orbital, followed by the s orbital and finally the d orbitals (if any more electrons need to be removed). For instance, the ground state electronic configuration of calcium (Z=20) is 1s²2s²2p⁶3s²3p⁶. The calcium ion (Ca²⁺), however, has two electrons less. Hence, the electron configuration for Ca²⁺ is 1s²2s²2p⁶3s²3p⁶. Since we need to take away two electrons, we first remove electrons from the outermost shell (n=4). In this case, all the 4p subshells are empty; hence, we start by removing from the s orbital, which is the 4s orbital. The electron configuration for Ca²⁺ is the same as that for Argon, which has 18 electrons. Hence, we can say that both are isoelectronic, having the same number of neutrons.

The electronic configuration of anions is assigned by adding electrons according to Aufbau’s building up principle. We add electrons to fill the outermost orbital that is occupied, and then add more electrons to the next higher orbital. The neutral atom chlorine (Z=17), for instance has 17 electrons. Therefore, its ground state electronic configuration can be written as 1s²2s²2p⁶3s²3p⁵. The chloride ion (Cl⁻), on the other hand, has an additional electron for a total of 18 electrons. Following Aufbau’s principle, the electron occupies the partially filled 3p subshell first, making the 3p orbital completely filled. The electronic configuration for Cl⁻ can, therefore, be designated as 1s²2s²2p⁶3s²3p⁶. Again, the electron configuration for the chloride ion is the same as that for Ca²⁺ and Argon. Hence, they are all isoelectronic to each other.

Brining It Full Circle

In Chapter 2, we discussed the charges of ions formed for main group elements as the gaining or losing of electrons to obtain the same number of electrons as the nearest noble gas. Now, we can thoroughly understand the reason for those charges using electron configurations. You have wondered why would calcium lose two electrons to form a Ca⁺² ion and be isoelectronic with Argon instead of gain 6 electrons to become Ca⁻⁶ and be isoelectronic with krypton. In terms of energetics, it takes much less energy to lose two electrons than to gain 6.

By the same token, chlorine will be isoelectronic with Argon if it gains one electron, but will have to lose seven electrons to be isoelectronic with neon. That is why, we see the charge of a chlorine ion as -1, not +7.
The same rule will apply to transition metals when forming ions. You should note that the ns electrons are always lost before the (n-1)d when forming cations for transition metals. For example,

the electron configuration for Zn: \[\text{[Ar]}4s^23d^{10}\]

the electron configuration for Zn\(^{+2}\): \[\text{[Ar]}3d^{10}\]

The transition metals still do not end up being isoelectronic with a noble gas, but the loss of two electrons is still beneficial due to achieving a more stable energy state for the system. However, how do we know that this is actually taking place and how do we trace what orbital(s) are losing/gaining electrons? We can study the magnetic properties of matter to help us tackle this problem.

The magnetic moment of a system measures the strength and the direction of its magnetism. The term itself usually refers to the magnetic dipole moment. Anything that is magnetic, like a bar magnet or a loop of electric current, has a magnetic moment. A magnetic moment is a vector quantity, with a magnitude and a direction. An electron has an electron magnetic dipole moment, generated by the electron's intrinsic spin property, making it an electric charge in motion. There are many different magnetic forms: including paramagnetism, and diamagnetism, ferromagnetism, and anti-ferromagnetism. Only paramagnetism, and diamagnetism are discussed here.

### Paramagnetism

Paramagnetism refers to the magnetic state of an atom with one or more unpaired electrons. The unpaired electrons are attracted by a magnetic field due to the electrons' magnetic dipole moments. Hund's Rule states that electrons must occupy every orbital singly before any orbital is doubly occupied. This may leave the atom with many unpaired electrons. Because unpaired electrons can orient in either direction, they exhibit magnetic moments that can align with a magnet. This capability allows paramagnetic atoms to be attracted to magnetic fields. Diatomic oxygen, \(\text{O}_2\) is a good example of paramagnetism (that is best understood with molecular orbital theory). The following video shows liquid oxygen attracted into a magnetic field created by a strong magnet

**Video 9.6.1: A chemical demonstration of the paramagnetism of molecular oxygen, as shown by the attraction of liquid oxygen to magnets.**

As shown in Video 9.6.1, since molecular oxygen \(\text{(O}_2\text{)}\) is paramagnetic, it is attracted to the magnet. In contrast, molecular nitrogen, \(\text{(N}_2\text{)}\), has no unpaired electrons and it is diamagnetic (discussed below); it is therefore unaffected by the magnet.

**Note**

Paramagnetism is a form of magnetism whereby materials are **attracted** by an externally applied magnetic field.

There are some exceptions to the paramagnetism rule; these concern some transition metals, in which the unpaired electron is not in a d-orbital. Examples of these metals include \(\text{(Sc}^{3+}\text{)}, \text{(Ti}^{4+}\text{)}, \text{(Zn}^{2+}\text{)}, \text{and (Cu}^{+}\text{).}

For example, you would expect the electron configuration of Cu to be: \(1s^22s^22p^63s^23p^64s^23d^9\) (paramagnetic, 1 unpaired
electron) and when it loses one electron to form the Cu⁺ with the following electron configuration: 1s²2s²2p⁶3s²3p⁶4s²3d⁸(paramagnetic; 2 unpaired electrons). In reality, the Cu⁺ ion is not attracted to a magnetic field, indicating that it has no unpaired electrons. This explains the anomalous electron configuration of the transition metals and allows us to refine the electron configuration of Cu as: 1s²2s²2p⁶3s²3p⁶4s¹3d¹⁰(paramagnetic, 1 unpaired electron) and so becomes Cu⁺: 1s²2s²2p⁶3s²3p⁶3d¹⁰(diamagnetic; no unpaired electrons) so that we are consistent with experimental data.

### Diamagnetism

Diamagnetic substances are characterized by paired electrons—except in the previously-discussed case of transition metals, there are no unpaired electrons. According to the **Pauli Exclusion Principle** which states that no two identical electrons may take up the same quantum state at the same time, the electron spins are oriented in opposite directions. This causes the magnetic fields of the electrons to cancel out; thus there is no net magnetic moment, and the atom cannot be attracted into a magnetic field. In fact, diamagnetic substances are weakly **repelled** by a magnetic field as demonstrated with the pyrolytic carbon sheet in Figure 9.6.1.

**Figure 9.6.1:** Levitating pyrolytic carbon: A small (~6 mm) piece of pyrolytic graphite levitating over a permanent neodymium magnet array (5 mm cubes on a piece of steel). Note that the poles of the magnets are aligned vertically and alternate (two with north facing up, and two with south facing up, diagonally). Image used with permission from Wikipedia.

Note

Diamagnetic materials are **repelled** by the applied magnetic field.
Diamagnetism, to a greater or lesser degree, is a property of all materials and always makes a weak contribution to the material's response to a magnetic field. For materials that show some other form of magnetism (such as paramagnetism), the diamagnetic contribution becomes negligible.

Practice

CHEMTOURS Tutorial: Click Chapter 7 “Electron Configuration”

How to tell if a substance is paramagnetic or diamagnetic

The magnetic form of a substance can be determined by examining its electron configuration: if it shows unpaired electrons, then the substance is paramagnetic; if all electrons are paired, the substance is diamagnetic. This process can be broken into four steps:

1. Find the electron configuration
2. Draw the valence orbitals
3. Look for unpaired electrons
4. Determine whether the substance is paramagnetic or diamagnetic

Example \(\PageIndex{1}\) : Chlorine Atoms

**Step 1: Find the electron configuration**

For Cl atoms, the electron configuration is \(3s^23p^5\)

**Step 2: Draw the valence orbitals**

Ignore the core electrons and focus on the valence electrons only.

![3s Orbital and 3p Orbitals](image)

**Step 3: Look for unpaired electrons**

There is one unpaired electron.
Step 4: Determine whether the substance is paramagnetic or diamagnetic

Since there is an unpaired electron, Cl atoms are paramagnetic (albeit, weakly).

Example \(\PageIndex{2}\) : Zinc Atoms

Step 1: Find the electron configuration

For Zn atoms, the electron configuration is 4s\(^2\)3d\(^{10}\)

Step 2: Draw the valence orbitals

\[
\begin{array}{c}
\uparrow \\
4s \text{ Orbital} \\
\hline \\
\hline \\
\uparrow \uparrow \uparrow \uparrow \uparrow \\
3d \text{ Orbitals}
\end{array}
\]

Step 3: Look for unpaired electrons

There are no unpaired electrons.

Step 4: Determine whether the substance is paramagnetic or diamagnetic

Because there are no unpaired electrons, Zn atoms are diamagnetic.

Exercise \(\PageIndex{1}\)

a. How many unpaired electrons are found in oxygen atoms?

b. How many unpaired electrons are found in bromine atoms?

c. Indicate whether boron atoms are paramagnetic or diamagnetic.

d. Indicate whether F\(^-\) ions are paramagnetic or diamagnetic.

e. Indicate whether Fe\(^{2+}\) ions are paramagnetic or diamagnetic.

Answer (a):

The O atom has 2s\(^2\)2p\(^4\) as the electron configuration. Therefore, O has 2 unpaired electrons.

Answer (b):

The Br atom has 4s\(^2\)3d\(^{10}\)4p\(^5\) as the electron configuration. Therefore, Br has 1 unpaired electron.

Answer (c):

The B atom has 2s\(^2\)2p\(^1\) as the electron configuration. Because it has one unpaired electron, it is paramagnetic.

Answer (d):

The F\(^-\) ion has 2s\(^2\)2p\(^6\) has the electron configuration. Because it has no unpaired electrons, it is diamagnetic.

Answer (e):

The Fe\(^{2+}\) ion has 3d\(^6\) has the electron configuration. Because it has 4 unpaired electrons, it is paramagnetic.
Contributors

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• Original Source 1

• Original Source 2