

1. Drawbacks/limitations of Bohr's atomic model.

➔ Certainly! Niels Bohr's atomic model, proposed in 1913, was a significant step in understanding the structure of atoms, but it also had several limitations. Here are some drawbacks of Bohr's atomic model along with their explanations:

1. Failure to Explain Line Spectra of Atoms with More Than One Electron:

- **Cause:** Bohr's model successfully explained the line spectra of hydrogen-like atoms with one electron, but it failed to predict the line spectra of atoms with more than one electron, such as helium or lithium. The model couldn't account for the complex patterns observed in their spectra.

2. Violation of Heisenberg's Uncertainty Principle:

- **Cause:** Bohr's model proposed definite orbits for electrons with well-defined positions and velocities. However, this violated the Heisenberg Uncertainty Principle, which states that it is impossible to simultaneously know the precise position and momentum of a particle. Bohr's model didn't conform to the inherent uncertainties at the microscopic level.

3. Inability to Explain the Zeeman Effect:

- **Cause:** The Zeeman effect, the splitting of spectral lines in the presence of a magnetic field, was not explained by Bohr's model. The model did not account for the interaction of electrons with magnetic fields, leading to inaccuracies in predicting experimental observations.

4. Fixed Orbits and Energy Levels:

- **Cause:** Bohr proposed that electrons orbit the nucleus in fixed, quantized orbits with specific energy levels. However, this concept contradicted the principles of classical mechanics and did not adequately explain the stability of these orbits. The model also failed to explain why electrons did not emit radiation while revolving in these stable orbits.

5. Neglect of Electron Spin:

- **Cause:** Bohr's model did not take into account the intrinsic spin of electrons, a property that is crucial for understanding the behavior of electrons in atoms. The inclusion of electron spin is essential for explaining phenomena like the fine structure of spectral lines.

6. Limited Applicability to Multielectron Atoms:

- **Cause:** Bohr's model primarily focused on hydrogen-like atoms, and its applicability diminished for atoms with more than one electron. The complexity of interactions between multiple electrons in different orbits was not addressed adequately.

2. Set of quantum numbers for 4p electron.

The four quantum numbers describe the unique properties of an electron in an atom. For a 4p electron:

1. Principal Quantum Number (n):

- For a 4p electron, $n = 4$. The principal quantum number represents the energy level or shell in which the electron is located.

2. Azimuthal Quantum Number (l):

- For a p orbital, $l = 1$. The azimuthal quantum number determines the shape of the orbital. For p orbitals, the possible values of l are $0, 1, 2, \dots, n-1$.

3. Magnetic Quantum Number (m_l):

- For a p orbital, m_l can take values from -1 to 1 , including 0 . This quantum number specifies the orientation of the orbital in space.

4. Spin Quantum Number (m_s):

- The spin quantum number describes the intrinsic angular momentum or "spin" of the electron. It has only two possible values: $+1/2$ or $-1/2$.

So, the set of quantum numbers for a 4p electron is:

$$n = 4,$$

$$l = 1,$$

m_l can be -1, 0, or 1,

m_s can be +1/2 or -1/2

3. Aufbau Principle. Explain exception along with electronic configuration.

The Aufbau Principle is a fundamental concept in quantum mechanics that describes the order in which electrons fill atomic orbitals in an atom. According to the Aufbau Principle:

1. Electrons fill the lowest energy orbitals first before moving to higher energy orbitals.
2. The order of filling is determined by the increasing energy levels (principal quantum number, n) and, within a given level, by increasing azimuthal quantum number (l).

Exceptions to the Aufbau Principle arise due to the fact that the energy levels of orbitals are affected by the electron-electron repulsion within a subshell. One notable exception involves the filling of the $4s$ and $3d$ orbitals.

For example, let's consider the electronic configuration of chromium (Cr, atomic number 24) and copper (Cu, atomic number 29):

1. Chromium (Cr):

- Expected configuration following Aufbau: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$
- Observed configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$

Chromium's observed configuration deviates from the expected order. Instead of filling $4s^2$ before $3d^4$, one electron is placed in the $4s$ orbital before

completing the 3d subshell. This irregularity is due to the stability gained when the 3d subshell is half-filled. A half-filled subshell is more stable than a partially filled 4s subshell.

2. Copper (Cu):

- Expected configuration following Aufbau:
 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
- Observed configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$

Similar to chromium, copper's observed configuration deviates from the expected order. Instead of $4s^2 3d^9$, one electron is placed in the 4s orbital before completing the 3d subshell. Again, this deviation is due to the increased stability of a fully filled 3d subshell.

In summary, the exceptions to the Aufbau Principle in the electronic configurations of chromium and copper occur because the stability gained from a half-filled or fully filled 3d subshell outweighs the expected order based solely on increasing energy levels.

4. Establish Hund's rule of maximum multiplicity.

Hund's Rule of Maximum Multiplicity is a principle in quantum mechanics that describes the filling of orbitals in an atom. It states:

1. Electrons occupy orbitals singly before pairing up.
2. When electrons occupy orbitals of equal energy (degenerate orbitals), they maximize the total spin quantum number (spin multiplicity).

To understand this rule better, let's consider the three p orbitals in a subshell (p_x , p_y , and p_z) as an example. These orbitals have the same energy. According to Hund's Rule:

- Electrons enter each orbital singly before pairing up.

- For the p orbitals, electrons would enter each (p_x , p_y , p_z) orbital with parallel spins (\uparrow) before any of them pair up with opposite spins ($\uparrow\downarrow$).

- Maximization of Spin Multiplicity.

- By having electrons with parallel spins in different orbitals, the total spin angular momentum is maximized. The term "spin multiplicity" refers to the number of possible spin states for a set of electrons. When electrons have parallel spins, the spin multiplicity is maximized.

This behavior arises from the fact that electrons are negatively charged particles, and they repel each other. By occupying separate orbitals with parallel spins, the electrons distribute themselves in a way that minimizes their repulsion, resulting in a lower overall energy for the atom.

Hund's Rule is not only applicable to p orbitals but also extends to other sets of degenerate orbitals, such as the d orbitals. The rule provides insight into the electron arrangement in atoms and is crucial for understanding the electronic structure of elements.

5. Increasing order of energy of 1st Bohr's orbit for H, He⁺, Li²⁺ and Be³⁺.

The energy levels of Bohr's orbits for hydrogen-like species are given by the formula:

$$E_n = -(Z^2 \cdot R_H) / n^2$$

where:

- E_n is the energy of the n th orbit,
- Z is the atomic number of the nucleus (number of protons),
- R_H is the Rydberg constant for hydrogen 2.18×10^{-18} J.
- n is the principal quantum number.

For the given species (H, He⁺, Li²⁺, Be³⁺), let's calculate the energy of the first Bohr orbit $n = 1$.

1. Hydrogen (H):

$$E_1(\text{H}) = -(1^2 \cdot R_H) / 1^2 = -R_H$$

2. Helium ion (He⁺):

$$E_1(\text{He}^+) = -(2^2 \cdot R_H) / 1^2 = -4 \cdot R_H$$

3. Lithium ion (Li²⁺):

$$E_1(\text{Li}^{2+}) = -(3^2 \cdot R_H) / 1^2 = -9 \cdot R_H$$

4. Beryllium ion (Be³⁺):

$$E_1(\text{Be}^{3+}) = -(4^2 \cdot R_H) / 1^2 = -16 \cdot R_H$$

Now, let's compare the energies in increasing order:

$$E_1(\text{H}) < E_1(\text{He}^{2+}) < E_1(\text{Li}^{2+}) < E_1(\text{Be}^{3+})$$

So, the increasing order of energy for the first Bohr orbit for the given species is:



6. Compare size of cation with corresponding atom(explain)

When a neutral atom loses one or more electrons to form a cation, it undergoes a change in its electronic configuration, which affects its size. The size of a cation is generally smaller than that of the corresponding neutral atom. Let's discuss the reasons behind this size difference:

1. Electron Removal:

- Cations are formed by the loss of one or more electrons from the outermost shell of an atom.
- The removal of electrons reduces the electron-electron repulsion within the electron cloud, making the remaining electrons experience a stronger effective nuclear charge.

2. Effective Nuclear Charge (Z_{eff}):

- The effective nuclear charge is the net positive charge experienced by an electron in an atom.
- When electrons are removed to form a cation, there are fewer electrons shielding the positive charge of the nucleus for the remaining electrons. As a result, the effective nuclear charge increases.

3. Size Contraction:

- The increase in effective nuclear charge pulls the remaining electrons closer to the nucleus.
- With fewer electron-electron repulsions and a stronger attractive force from the nucleus, the electron cloud contracts, leading to a reduction in the size of the cation compared to the neutral atom.

4. Example: Sodium (Na) and Sodium Ion (Na⁺):

- Sodium has an electron configuration of $1s^2 2s^2 2p^6 3s^1$.
- When it loses its outermost electron to become a sodium ion (Na⁺), the electron configuration becomes $1s^2 2s^2 2p^6$.
- The effective nuclear charge increases, and the electron cloud contracts, resulting in a smaller size for the Na⁺ ion compared to the neutral Na atom.

In summary, the size of a cation is smaller than the corresponding neutral atom due to the removal of electrons, an increase in effective nuclear charge, and the resulting contraction of the electron cloud. This trend holds true for most cations across the periodic table.

7. Size of atom, Ionization energy along group and across period

The size of an atom and its ionization energy are influenced by atomic structure and can be understood by considering trends along a group (column) and across a period (row) in the periodic table.

1. Size of Atom:

- Along a Group:

- Down a group, the size of atoms increases. This is because each successive element in a group has an additional energy level (shell), leading to an increase in the average distance of electrons from the nucleus. The outermost electrons are in higher energy levels, and the inner electrons shield the outer electrons from the attractive force of the nucleus, allowing the electron cloud to expand.

- Across a Period:

- Across a period, the size of atoms generally decreases. This is due to increasing effective nuclear charge as electrons are added to the same energy level. The additional electrons experience a greater attraction to the nucleus, resulting in a more compact electron cloud.

2. Ionization Energy:

- Along a Group:

- Down a group, ionization energy tends to decrease. This is because electrons are farther from the nucleus in higher energy levels, and the outer electrons are shielded by inner electrons. As a result, it is easier to remove an electron.

- **Across a Period:**

- Across a period, ionization energy generally increases. This is because the effective nuclear charge increases (more protons in the nucleus), leading to a stronger attraction between electrons and the nucleus. As a result, more energy is required to remove an electron.

In summary:

- **Size of Atom:** Increases down a group, decreases across a period.

- **Ionization Energy:** Decreases down a group, increases across a period.

These trends are generalizations and may have exceptions due to factors such as electron-electron repulsions and subshell configurations. Additionally, there are anomalies in ionization energy caused by stable half-filled and fully filled subshells, which can result in slightly lower ionization energies for elements in certain groups.

8. Compare Electron affinity of F and Cl with explanation.

Electron affinity is the energy change that occurs when an atom gains an electron to form a negative ion (anion). Generally, electron affinity tends to increase across a period from left to right and decrease down a group in the periodic table.

Let's compare the electron affinities of fluorine (F) and chlorine (Cl):

1. Fluorine (F):

- Fluorine is in Group 17 (Group VIIA), and it is close to the right end of the period.
- It has a high electron affinity because it needs one more electron to achieve a stable, noble gas electron configuration.
- The addition of an electron to fluorine releases energy, and the process is exothermic.

2. Chlorine (Cl):

- Chlorine is also in Group 17, but it is one step below fluorine.
- Like fluorine, chlorine has a high electron affinity, but it is slightly lower than that of fluorine.
- The additional electron is added to a higher energy level in chlorine compared to fluorine, and the increased distance from the nucleus reduces the attraction, making the process slightly less exothermic.

In summary, both fluorine and chlorine have relatively high electron affinities due to their position in Group 17, indicating a strong tendency to gain an electron. However, fluorine has a slightly higher electron affinity than chlorine. The trend of increasing electron affinity across a period from left to right is generally observed due to the increasing effective nuclear charge, which enhances the attraction between the nucleus and the incoming electron.

9. Compare ionization potential for Nitrogen and Oxygen with explanation.

Ionization potential, also known as ionization energy, is the energy required to remove an electron from an atom or a positive ion. It is usually measured in units of energy per mole, such as kilojoules per mole (kJ/mol) or electron volts (eV). Let's compare the ionization potentials of nitrogen (N) and oxygen (O):

1. Nitrogen (N):

- Nitrogen is in Group 15 of the periodic table and has an electron configuration of $1s^2 2s^2 2p^3$.
- The ionization potential of nitrogen is relatively low compared to elements in later periods because nitrogen has a half-filled 2p subshell. Half-filled and fully filled subshells tend to have greater stability, making it easier to remove an electron from the outermost shell.

2. Oxygen (O):

- Oxygen is in Group 16 and has an electron configuration of $1s^2 2s^2 2p^4$.
- The ionization potential of oxygen is higher than that of nitrogen because removing an electron involves breaking into the more stable half-filled 2p subshell. Oxygen has one additional electron in the 2p subshell compared to nitrogen, and the extra electron increases the electron-electron repulsion, making it slightly more difficult to remove an electron.

In summary, the ionization potential of nitrogen is lower than that of oxygen. This is mainly attributed to the favorable stability of the half-filled 2p subshell in nitrogen, which makes it easier to remove an electron compared to oxygen with its additional electron in the 2p subshell.