How Does the Screening Effect Influence Atomic Properties?

The **screening effect** is also known as the shielding effect. In the middle of each atom is a nucleus, surrounding which the electrons will circle in various orbits. We are aware that the outer electrons are drawn to and bound in an atom by nuclear charges. The inner shell electrons shield the outermost shell electrons from the nuclear force. The valence electron is the electron in an atom's outermost shell.

The **screening effect** governs the atom's ionisation enthalpy. The amount of energy needed to remove an electron from a free gaseous atom is known as the ionisation enthalpy. The nuclear force and ionisation energy are inversely correlated because a strong nuclear force attracts its valence electrons, which increases ionisation energy.

Therefore, it will effectively bind electrons, making it challenging to release them. So, the energy needed for removal will be higher. The ionisation energy needed will be low if the **screening effect** reduces nuclear force. Therefore, the **screening effect** will cause the ionisation energy to decline. Thus, this article provides in-depth information about **screening effects in Chemistry**.

What is the Screening Effect in Chemistry?

The **screening effect in Chemistry** is known as a decrease in the effective nuclear charges on the electron cloud as a result of a variation in the force of attraction acting on the atom's electrons. This particular instance of **screening effect** is unique. This **screening effect** is also important in a number of material science initiatives.

The screening effect of inner electrons of the nucleus causes repulsion, and the **screening effect** of outermost electrons of the nucleus causes attraction. As the attraction connecting the nuclei weakens, the repulsion involving the outer and inner electrons intensifies. Owing to the electrostatic force exerted by the protons in the nucleus, the outermost electron is attracted to the nucleus.

Order of Screening Effect in Electron Shells

Owing to the **screening effect**, the electric connection involving the electrons and the nucleus is poorer; when broader the electron shells are spread out in space. In general, the **screening effect** in shells (s, p, d, and f) is as follows:

s > p > d > f

There are two separate forces acting on the outer shell electrons. The initial one is the nucleus's attraction, and the latter one is the inner shell electrons' repulsion.

Screening Effect Description

The force on the electron in hydrogen, or some similar element in 1A group of the periodic table (elements with just single outer shell electrons), is equal to the electromagnetic force on the atom’s nuclei. Whenever additional electrons are engaged, every electron (inside the nth shell) encounters not just the electromagnetic force from the positively charged nucleus but also repulsive forces by additional electrons in shells numbered 1 to n.

As a result, the overall force on electrons within the outermost shells is much weaker in magnitude; hence, these electrons cannot be as tightly linked to the nuclei with electrons nearer to the nuclei. This process is known as the orbital penetrating effect. The **screening effect** also helps to explain why outer electrons are easier to remove from the atom.

Among sub-levels of the equivalent major energy level, there is furthermore a **screening effect**. An electron at a similar fundamental energy level's s-sublevel has the ability to protect electrons at the p-sublevel. This is due to the s-spherical orbital structure. The opposite is untrue: electrons in an s-orbital cannot be shielded by electrons from a p-orbital.

Screening Effect with Example

Since the electrons are introduced to the same shell throughout the period, the **screening effect** caused by inner electrons does not change. As a fresh valence shell is introduced, the **screening effect** caused by inner electrons rises down the group.

The electronic configuration of potassium (K) is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

Since potassium contains four shells, the electrons inside the inner 3 shells serve as a shield for the outermost shell electrons. This causes the outermost shell electron $4s^1$ in potassium to have a significantly weaker nuclear charge and can be eliminated with ease.

Interesting Facts

- The shielding effect is greater as we travel the group downwards in the periodic table than the influence of increasing protons. As the ionisation energy of Alkalis declines down the group, the valence electrons in the s-shell are rapidly lost.
- The electron in the s-shell covers or screens the electrons in the p-shell. Additionally, the d-shell electrons are screened by the electrons of the s and p shells.
- The **screening effect** of the inner shell electrons protects the valence electrons against the nuclear force of attraction. The desired ionisation enthalpy increases with the nuclear force of attraction.
- The **screening effect** has an inverse relationship with ionisation enthalpy.

Key Features to Remember

- The electrons in an atom with more than one electron prefer to keep the valence electrons from being attracted to the nucleus by the inner shell electron.
- Thus the **screening effect** occurs when the nucleus's attractive attraction weakens relative to the inner shell electrons' repulsive force.
- The electron cloud's apparent nuclear charge is decreased as a result of shielding. The "shielding effect" is the other name for the **screening effect**.
- The broader the electron shells are in space, the poorer the screening causes the electrical bonding involving the electrons and the nucleus.