

## Quadrant II – Transcript and Related Materials

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### ELECTRON AFFINITY

Electron Affinity is defined as **“the amount of energy released when a gaseous isolated atom accepts an extra electron to form the gaseous negative ion.”**

Electron Affinity is abbreviated as E.A. and expressed in terms of electron volts per atom or kJ per mol.

This can be expressed by an equation as ...  $A(g) + e^{-} \rightarrow A^{-}(g) + \text{Energy}$

Since energy is released E.A. is a negative quantity; although positive values are also observed but for ionic species with -2 charge like  $O^{2-}$  and  $S^{2-}$  where the E.A. in this case is referred to as “the second electron affinity”.

**Second electron affinity is defined as “the energy required to add an electron to each ion in one mole of gaseous 1- ions to produce one mole of gaseous 2- ions.”**



The positive sign indicates that energy has to be supplied to perform the change.

E.A. exhibits varying trends across a period and down the group in the Periodic Table. There are three main factors responsible for these variations in trends observed here, namely, Atomic Size, Nuclear Charge and Electron configuration.

- **ATOMIC SIZE:** Smaller the size of the atom, lesser will be the distance between the nucleus and the added extra electron. Therefore greater will be the electrostatic force of attraction and greater will be the E.A.

- **NUCLEAR CHARGE:** Higher the nuclear charge, greater will be the attraction or affinity for the additional electron and therefore, greater will be the electron affinity.
- **ELECTRON CONFIGURATION:** Electron configuration can predict the stability of an atom. Elements with half-filled or fully-filled orbitals are very stable and so will not attract additional electrons. Thus they will exhibit zero or very low Electron affinities.

**The trends for E.A. observed in the Periodic Table are as follows:**

Across a Period: When we move from left to right, across the Periodic Table, the E.A. increases

Down the Group: When we move from top to bottom, the E.A. decreases.

**Across the Period:** As we move across the period, the Atomic Size decreases. So the outermost orbital will be closer to the nucleus. Naturally, therefore, the Nuclear Charge increases. As a result, there is increased attraction for the additional electron by the nucleus, hence greater energy will be released, which means increase in electron affinity.

**Down the Group:** As we move down the Group, the Atomic Size and the Nuclear Charge increase. At the same time there is also increase in the number of electron shells. This outweighs the increase in the nuclear charge, as a result of which the effective Nuclear Charge decreases. Hence the atom will have a lesser tendency to attract additional electron towards itself. Consequently lesser energy will be released, which means decrease in electron affinity.

Although these are the normal trends, **some exceptions or anomalies are observed** (highlighted in yellow), which are due to the special stabilities of orbitals.

**Examples :** Elements of **Group 2 A** having outermost electronic configuration as  $ns^2$  as well as **noble gases** having outer electronic configuration as  $ns^2np^6$  (both fully filled outer orbitals or stable octet in the case of noble gases); and **Nitrogen** with outer electronic configuration of  $ns^2, np^3$  (half-filled stable p-orbitals) show 0 or very low electron affinities, because in these cases since the orbitals are already stable, adding an electron will disturb their stability and instead of an electron being taken up easily it will have to be forced into these orbitals by supplying energy.

In this way, we can explain the position of a particular element in the periodic table and anomalies observed as well.

HAPPY LEARNING!