Octet Rule Encyclopedia Article

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BOOKRAGS

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The octet rule is used to describe the attraction of elements towards having, whenever possible, eight valence-shell electrons (4 **electron** pairs) in their outer shell. Because a full outer shell with eight electrons is relatively stable, many atoms lose or gain electrons to obtain an electron configuration like that of the nearest noble gas. Except for **Helium** (with a filled 1s shell), noble **gases** have eight electrons in their **valence** shells. The octet rule is used when drawing Lewis dot structures.

The components of table **salt**, **sodium** and chloride ions, illustrate the octet rule. Sodium (Na) with an electron configuration of $1s^2 2s^2 2p^6 3s^1$ sheds its outermost 3*s* electron and, as a result, the Na^+**ion** has an electron configuration of $1s^2 2s^2 2p^6$. This is the same electron configuration as **neon**, a noble gas (i.e., highly stable and relatively nonreactive).

Chlorine (Cl), on the other hand, has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^5$. Chlorine needs one electron to fill its outermost third shell with eight electrons. When chlorine takes on the electron shed by sodium then the Cl^+ ion's electron configuration becomes $1s^2 2s^2 2p^6 3s^2 3p^6$. This is the same configuration as **Argon**, another noble gas.

The octet rule, however, does not accurately predict the electron configurations of all molecules and compounds. Not every **nonmetal**, nor metal, can form compounds that satisfies the octet rule. As a result, the octet rule must be used with caution when predicting the electron configurations of molecules and compounds. Some atoms violate the octet rule and surround themselves with more than four electron pairs.

In general, the octet rule works for representative **metals** (Groups IA, IIA) and nonmetals, but not for the transition, inner-transition or post-transition elements. These elements seek additional stability by having filled half-filled or filled orbitals *d* or *f* subshell orbitals.