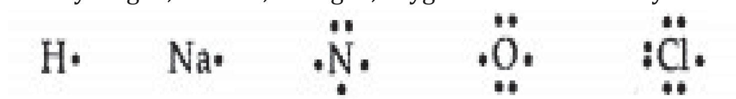
LEWIS DOT STRUCTURE

Lewis Symbols of Elements

Chemical bonding primarily relies on the count of electrons found in the outermost energy level, commonly referred to as valency electrons. Sodium (Na) exhibits an electronic configuration of 2, 8, 1, signifying one valency electron, while sulfur (S) has an electronic configuration of 2, 8, 6, indicating six valency electrons. For representative elements, the group number in the Modern Mendeleev's periodic table corresponds to the number of valency electrons.

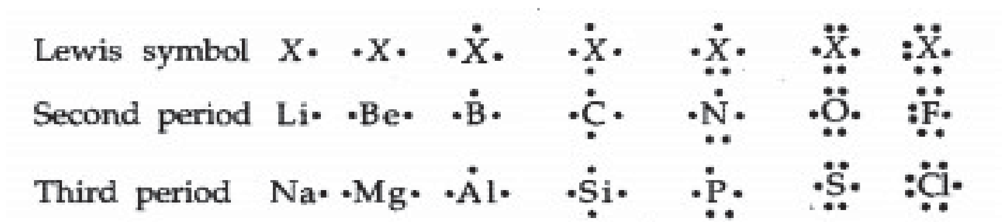
The depiction of valency electrons in atoms is represented using Lewis symbols. These symbols are constructed by placing the element's symbol surrounded by a specific number of dots or crosses, corresponding to the count of valency electrons. Additionally, paired and unpaired valency electrons are delineated in the Lewis symbols.

The Lewis symbols for hydrogen, sodium, nitrogen, oxygen and chlorine may be written as:



Generalized, Lewis symbols for the representative elements are given in the following table:

	1	2	13	14	15	16	17
Group	IA	IIA	IILA	IVA	VA	VIA	VIIA
	ns^1	ns^2	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5



Lewis Octet Rule

The Octet Rule, formulated by Lewis and Kossel, asserts that atoms strive to achieve a stable state by combining to complete an octet of electrons in their outermost orbit. This octet, represented by the s^2p^6 configuration, signifies maximum stability. Consequently, all atoms exhibit a tendency to acquire this octet configuration in their outermost orbit.

The attainment of an octet can occur through various means, including the complete transfer of electrons from one atom to another.

Ex. NaCl, $CaCl_2$ & MgO etc. (Ionic Bond) Sharing of electrons between atoms.

(a) Sharing of equal number of electrons between two atoms.

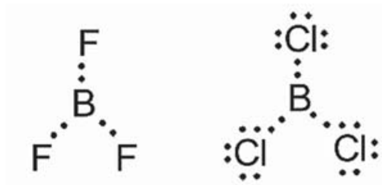
Ex. Cl_2, N_2, O_2 etc., (Covalent bond)

(b) Sharing of electron pair given by only one atom

Ex. [$NH_3 \rightarrow H^+$] & $NH_3 \rightarrow BF_3$ (Co-ordinate Bond)

Limitations of Lewis Octet Rule

(i) **Incomplete Octet of the Central Atom:** Certain compounds deviate from the conventional octet rule as the central atom may not possess a complete octet but remains stable nonetheless.



(ii) **Expanded Octet Rule:** Elements in the 3rd period, such as sulfur hexafluoride (SF_6) and phosphorus pentachloride (PCl_5), challenge the eight-electron norm, demonstrating the existence of compounds with central atoms that exceed the traditional octet, featuring 12 and 10 valence electrons, respectively.

(iii) **Noble Gas Xenon (Xe) Forming Compounds:** Xenon, typically a noble gas, defies expectations by forming compounds, challenging the conventional understanding of noble gases as inert and unreactive.

(iv) **Molecular Shape Unaddressed:** Lewis's octet rule falls short in providing insights into the three-dimensional molecular shape of a compound, omitting valuable information about the spatial arrangement of atoms.

While Lewis dot representations offer a valuable depiction of bonding through shared pairs of electrons and adherence to the octet rule, they do not offer a comprehensive understanding of the complete bonding picture. Despite these limitations, Lewis's dot structures contribute to the comprehension of molecular formation.

Exceptions of Octet Rule

1. **Transition Metal Ions**

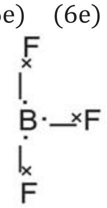
Cr^{3+}	Mn^{2+}	Fe^{2+}
$[Ar]3d^3$	$[Ar]3d^5$	$[Ar]3d^6$
[2, 8, 11]	[2, 8, 13]	[2, 8, 14]

2. **Pseudo Inert Gas Configuration $[s^2p^6d^{10}]$**

Zn^{2+}	Cd^{2+}
$[Ar]3d^{10}$	$[Kr]4d^{10}$

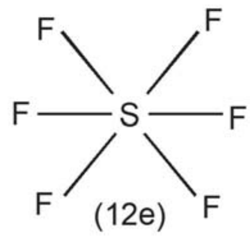
3. **Contraction of Octet (Incomplete Octet)**

BeF_2	BF_3	$AlCl_3$	BCl_3
(4e)	(6e)	(6e)	(6e)



4. **Expansion of Octet (Due to Empty d-Orbitals)**

PCl_5	SF_6	ClF_3	ICl_5	IF_7
↓	↓	↓	↓	↓
(10e)	(12e)	(10e)	(12e)	(14e)



Formal Charge

The molecule under consideration is electrically neutral, signifying that its constituent atoms lack any charge. This characteristic is applicable not only to individual atoms but also extends to polyatomic ions, where the overall charge is attributed to the ion as a collective entity rather than any specific atom within it. Despite this overarching principle, there are situations where it proves advantageous to designate formal charges to the individual atoms forming molecules or polyatomic ions.

Formal charge, in the context of an atom, is determined by assessing the disparity between the number of valence electrons in an isolated atom (i.e., a free atom) and the number of electrons ascribed to that atom within a Lewis structure.

The formal charge is articulated by the formula:

Formal Charge = [Number of Valence Electrons in an Isolated Atom] – [Number of Electrons Assigned in the Lewis Structure]

This computation is grounded on the assumption that each shared pair of electrons contributes one electron to the atom in the molecule, and both electrons of a lone pair are considered.

To illustrate, let's examine the H_3PO_4 molecule. In this particular molecule, the formal charge on the phosphorus (P) atom can be calculated utilizing the aforementioned formula.

- The central phosphorus atom (P) has five valence electrons.
- Each oxygen atom (O) has six valence electrons, but two of these are lone pairs and do not participate in bonding.
- The oxygen atoms are bonded to the phosphorus atom by single covalent bonds, using two of their lone pairs.
- Each hydrogen atom (H) has one valence electron, which it shares with an oxygen atom to form a single covalent bond.

The formal charge on an atom in a molecule is the difference between the number of valence electrons it has in the free state and the number of bonding electrons it has in the molecule plus the number of lone pairs it has. In the HOPOH molecule, the phosphorus atom has five valence electrons, and it shares two of these with each oxygen atom, for a total of four bonding electrons. It also has one lone pair.

Therefore, the formal charge on the phosphorus atom is 5 - 4 - 1 = 0.

The oxygen atoms each have six valence electrons, but two of these are lone pairs. They share two of their remaining four valence electrons with the phosphorus atom to form bonding electrons.

Therefore, each oxygen atom has a formal charge of 6 - 2 - 2 = 2.