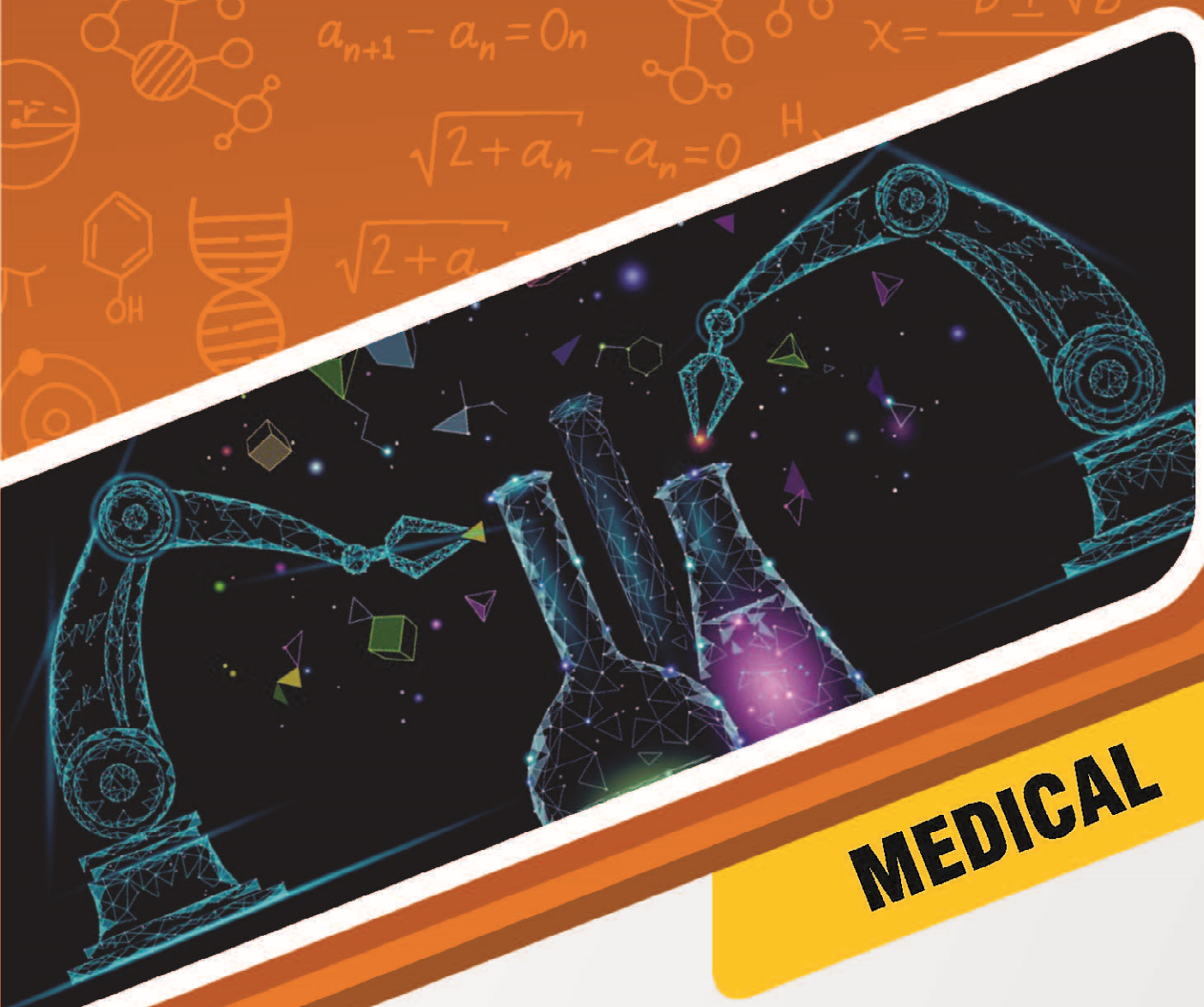
CHEMISTRY



MEDICAL

CHEMICAL BONDING

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CONTENT

CHEMICAL BONDING

Chemical Bonding

INTRODUCTION

(i) Most of the elements exist as molecules which are cluster of atoms. How do atoms combine to form molecules and why do atoms form bonds. Such doubts will be discussed in this chapter.

(ii) A molecule will only be formed if it is more stable and has a lower energy, than the individual atoms.

CHEMICAL BOND

(i) A force that acts between two or more atoms to hold them together as a stable molecule.

(ii) It is union of two or more atoms involving redistribution of electron among them.

(iii) This process accompanied by decrease in energy.

(iv) Decrease in energy $\propto$ Strength of the bond.

(v) Therefore molecules are more stable than atoms.

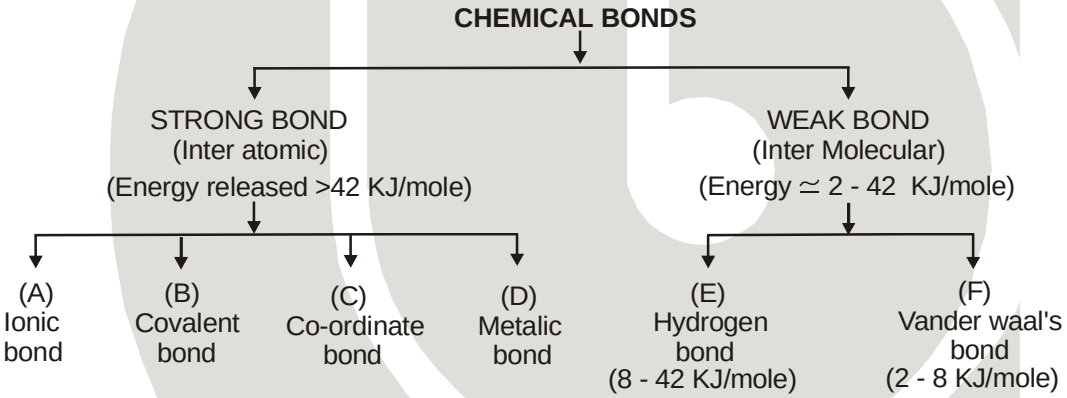
CAUSE OF CHEMICAL COMBINATION

1. Tendency to acquire noble gas configuration : (classical concept)

(i) Atom combines to acquire noble gas configuration.

(ii) Only outermost electron i.e. ns, np and (n-1)d electrons participate in bond formation.

(iii) Inert gas elements do not participate, as they have stable electronic configuration and hence minimum energy. (Stable electronic configuration : $1s^2$ or ns^2np^6)



2. Tendency to acquire minimum energy :(modern concept)

(i) When two atoms approaches to each other. Nucleus of one atom attracts the electron of another atom.

(ii) Two nuclei and electron of both the atoms repells each other.

(iii) If net result is attraction, the total energy of the system (molecule) decreases and a chemical bond forms.

(iv) So, Attraction $\propto$ 1/energy $\propto$ Stability.

1. OCTET RULE (LEWIS-KOSSEL RULE) :

Noble gases have 8 electron in their outermost shell (complete octet) and outermost configuration is ns^2p^6. Every atom has a tendency to complete its octet by losing or gaining or by sharing electron.

Three different types of bonds—ionic, covalent, and metallic—are formed depending on the electropositive or electronegative character of the atoms (of an element) involved in a bond formation. Atoms of various elements may be electropositive or electronegative depending on the type of element. Elements may be classified as follows:

(i) Electropositive elements. These elements readily lose one or more electrons.
(ii) Electronegative elements. These elements accept electrons and achieve a stable electronic configuration.
(iii) Elements that have a little tendency to lose or gain electrons.

ELECTRON-DOT (LEWIS) STRUCTURE

For single central atom Lewis structure drawing

Rules for Lewis Structure Drawing :

(i) Calculate n_1 = no. of valence shell electrons of all atoms
+ no. of negative charge (if any)
– no. of positive charge (if any)

(ii) Calculate n_2 = (no. of H-atom × 2) + (no. of atoms other than H-atoms × 8)

(iii) Calculate $n_3 = n_2 - n_1 \Rightarrow$ no. of shared electrons

i.e. No. of bonds = $\frac{n_3}{2}$ .....(a)

(iv) Calculate $n_4 = n_1 - n_3 \Rightarrow$ no. of unshared electrons

i.e. no. of lone pairs = $\frac{n_4}{2}$ .....(b)

Using the informations (a) & (b) the structure is to be assigned as follows:

(i) Find out the central atom first (i.e. either least in number or more electro positive)
(ii) Allocate the surrounding atoms around the central atom with the help of bonds available in (a).
(iii) To fulfill the octet of each atom. utilise the lone pairs available in (b).
(iv) Finally calculate the formal charge for each atom and assign on the atoms according the formula given.

F.C. (of an atom) = Valence shell electron of that atom
– no. of bonds associated with it
– no. of unshared electrons on it.

In order to write the Lewis structure of a molecule, the points given below need to be kept in mind.

(i) Add the valence electrons of all atoms. In case of an ion, add electrons equal to negative charge (for negative ion) or subtract the number of electrons equal to the positive charge (for a positive ion).

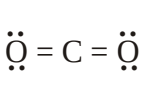
(ii) Write the formula of the compound, and join atoms by single bonds.

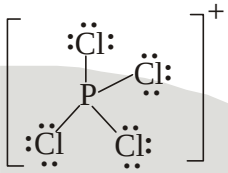
(iii) Complete the octet of all the atoms (or two electrons for hydrogen) which are bonded to the central atom.

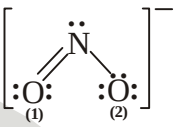
(iv) Fix up the remaining electrons, if any, on the central atom.

(v) In case the central atom does not have the requisite number of electrons to complete its octet, attach it by multiple bonds.

S.NO.		CO_2	NH_3	PCl_4^+	NO_2^-
1.	n_1	16	8	32	18
2.	n_2	24	14	40	24
3.	$n_2 - n_1 = n_3$	8	6	8	6
4.	$n_3/2$	4	3	4	3
5.	$n_3 - n_1 = n_4$	8	2	24	12
6.	$n_4/2$	4	1	12	6







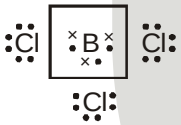
Formal change	C = zero	N = zero	P = +1	N = zero
	O = zero	H = zero	Cl = zero	O(1) = zero
				O(2) = –1

1 EXCEPTIONS OF OCTET RULE

(a) Incomplete octet molecules : or (electron defficient molecules) or Hypovalent molecules

Compound in which octet is not complete in outer most orbit of central atom.

Examples - Halides of IIIA groups, BF_3, $AlCl_3$, BCl_3, hydride of III A/13th group etc.



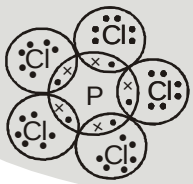
In BCl_3
Boron has only 6 electrons

Other examples - $BeCl_2$ (4 electron), $Ga(CH_3)_3$ (6 electron)

(b) Expansion of octet or (electron efficient molecules) or Hypervalent molecules

Compound in which central atom has more than 8 electron in outermost orbits.

Ex. PCl_5, SF_6, IF_7, the central atom P, S and I contain 10, 12, and 14 electrons respectively.



Electron dot formula of PCl_5

(c) Pseudo inert gas configuration :

Cations of transition metals, which contains 18 electrons in outermost orbit

Examples : Ga^{+3}, Cu^+, Ag^+, Zn^{+2}, Cd^{+2}, Sn^{+4}, Pb^{+4} etc.

Electronic configuration of Ga - $1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^1$