
GENERAL CHEMISTRY
4
Dr M.F Dr A.E

Chapter Three

Part (1)

✓ Ionic Bond

Chemical Bonds

Force at which atoms are held together.

Valence Electrons

Electrons of the outermost shell (responsible for bonding)

Note Be

Valence electron = Group Number

Group	Outershell	No fo valence e's
gp I (1A)	ns^1	1
gp II (2A)	ns^2	2
gp III (3A)	ns^2np^1	3
gp IV (4A)	ns^2np^2	4
gp V (5A)	ns^2np^3	5
gp VI (6A)	ns^2np^4	6
gp VII (7A)	ns^2np^5	7
gp VIII (8A)	ns^2np^6	8

Octet Rule

Atoms tend to gain, lose or share e's to be surrounded by 8 e's in the outer shell to have the same electronic configuration of noble gas. (To Be Stable)

✔ Except Hydrogen gas → surrounded by 2 electrons

Lewis Dot symbol for atoms

1. $_{8}O$ → gp 6 → ∴ Valence electrons = 6 $\cdot\ddot{\underset{\cdot\cdot}{O}}\cdot$

2. $_{9}F$ → gp 7 → ∴ Valence electrons = 7 $:\ddot{\underset{\cdot\cdot}{F}}\cdot$

3. $_{16}S$ → gp 6 → ∴ 6 Valence e's $\cdot\ddot{\underset{\cdot\cdot}{S}}\cdot$

4. $_{16}S^{2-}$ → gp 6 → ∴ 6 Valence e's + 2 gain e's $:\ddot{\underset{\cdot\cdot}{S}}:$

Types of Chemical Bonds

1) Ionic Bond

Electrostatic attraction between +Ve ion and -Ve ion

Between Metal And Non-Metal

gp IA	gp IIA	gp IIIA	gp VA	gp VIA	gp VIIA
Li^{+}	Mg^{2+}	Al^{3+}	N^{3-}	O^{2-}	F^{-}
Na^{+}	Ca^{2+}		P^{3-}	S^{2-}	Cl^{-}
K^{+}	Sr^{2+}				Br^{-}
Rb^{+}	Ba^{2+}				I^{-}
Cs^{+}					

Involve transfer of electrons From Metal to Non-Metal

Metal	Non-Metal
M (low IE)	X (high EA)
Loss e's	Gain e's
M^{+} Cation	X^{-} Anion

MX (Ionic Compound)

Example

1] Formation of NaCl

$$Na\cdot \quad + \quad \cdot\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}: \quad \rightarrow \quad Na^{+} \; [:\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}:]^{-}$$

(low IE) (high EA)

✔ $_{11}Na^{+}$ ⟶ $1S^2\ 2S^2\ 2P^6\ 3S^0$ Isoelectronic with Neon

✔ $_{17}Cl^{-}$ ⟶ $1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^6$ Isoelectronic with Ar

2] Formation of MgO

$$Mg: \quad + \quad \cdot\underset{\cdot\cdot}{\overset{\cdot\cdot}{O}}\cdot \quad \rightarrow \quad Mg^{2+} \; [:\underset{\cdot\cdot}{\overset{\cdot\cdot}{O}}:]^{2-}$$

(low IE) (high EA)

✔ $_{12}Mg^{2+}$ ⟶ $1S^2\ 2S^2\ 2P^6\ 3S^0$ Isoelectronic with Neon

✔ $_{8}O^{2-}$ ⟶ $1S^2\ 2S^2\ 2P^6$ Isoelectronic with Neon

♦ Lattice Energy (U_L)

The energy required to separate completely a mole of solid ionic compound into its ions in gaseous state

$NaCl_{(s)} \rightarrow Na^{+}_{(g)} + Cl^{-}_{(g)}$ ∴ $U_L = +788$ KJ (required)

Or

The energy released for formation of one mole of solid crystalline compound from its ions in gaseous state

$Na^{+}_{(g)} + Cl^{-}_{(g)} \rightarrow NaCl_{(s)}$ ∴ $U_L = -788$ KJ (released)

$$U_L = K\frac{Q_1 Q_2}{d}$$

❖ Q_1 and Q_2 : Electric Charge Of ions

❖ d : Distance between ions ❖ K : Constant

Important Notes

OR
❖ As Charge ↑
❖ As Size ↓

- Bond Strength ↑
- Stability ↑
- Melting point (m.p) ↑
- Lattice Energy ↑
- Solubility ↓

Ex: Arrangement according to stability, m.p and Lattice Energy

1- $Na^+F^- > Na^+Cl^- > Na^+Br^- > Na^+I^-$

2- $Li^+Cl^- > Na^+Cl^- > K^+Cl^- > Rb^+Cl^- > Cs^+Cl^-$

3- $Al^{3+}Cl_3^- > Mg^{2+}Cl_2^- > Na^+Cl^-$

4- $Cs^+F^- > Cs^+Cl^- > Cs^+Br^- > Cs^+I^-$

5- $Mg^{2+}O^{2-} > Ca^{2+}O^{2-} > Mg^{2+}Cl_2^-$

Ex: Arrangement the ionic compounds according to Solubility:

1) $Na^+F^- < Na^+Cl^- < Na^+Br^- < Na^+I^-$

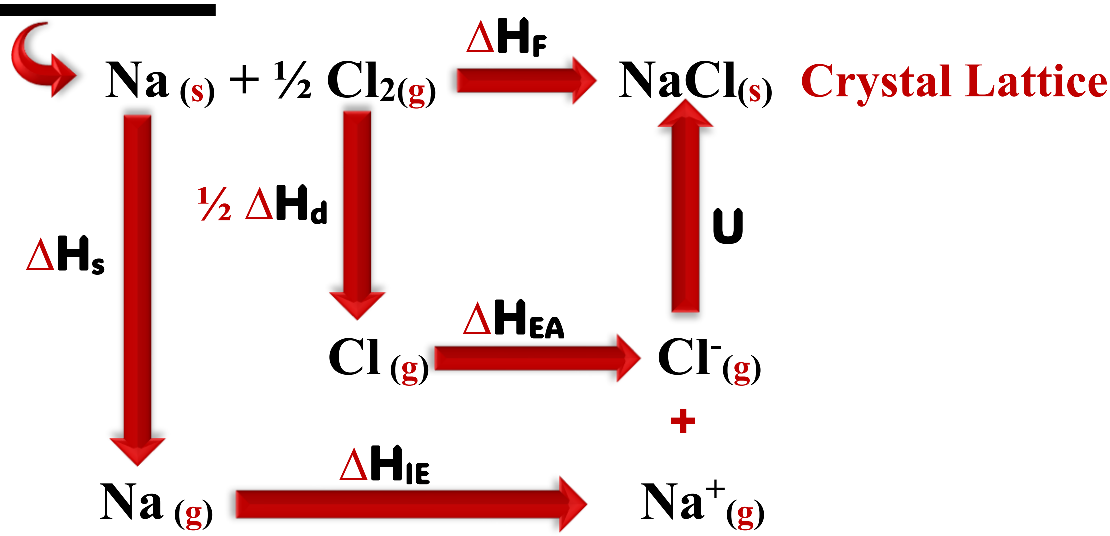
2) $Li^+Cl^- < Na^+Cl^- < K^+Cl^- < Rb^+Cl^- < Cs^+Cl^-$

3) $Al^{3+}Cl_3^- < Mg^{2+}Cl_2^- < Na^+Cl^-$

♦ Formation Of ionic Crystal (Born Haber Cycle)

♦ **Based on** (Hess Law)

♦ **Ex : NaCl**



$$\boxed{\Delta H_F = \Delta H_s + ½\,\Delta H_d + \Delta H_{IE} + \Delta H_{EA} + U}$$

- ΔH_F : Heat of formation
- ΔH_s : Heat of Sublimation
- ΔH_d : Heat of dissociation
- ΔH_{IE} : Ionization Energy
- ΔH_{EA} : Electron affinity
- U : Lattice energy

Ex: Calculate lattice energy for the following reaction:

1- $Na_{(s)} + ½\,Cl_{2(g)} \rightarrow NaCl_{(s)}$ $\Delta H_f = -\ 411$ KJ/mol

2- $Na_{(s)} \rightarrow Na_{(g)}$ $\Delta H_s = +\ 108$ KJ/mol

3- $Na_{(g)} \rightarrow Na^+_{(g)} + e^-$ $\Delta H_{IE} = +\ 500$ KJ/mol

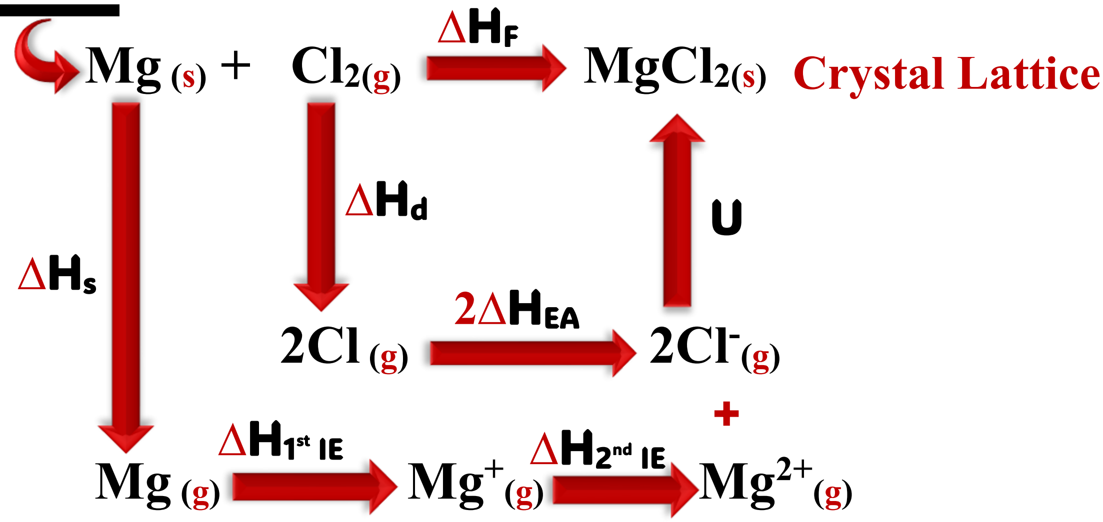
4- $½\,Cl_{2(g)} \rightarrow Cl_{(g)}$ $½\,\Delta H_d = -121.1$ KJ/mol

5- $Cl_{(g)} + e^- \rightarrow Cl^-_{(g)}$ $\Delta H_{EA} = -349$ KJ/mol

∴ $\Delta H_F = \Delta H_s + ½\,\Delta H_d + \Delta H_{IE} + \Delta H_{EA} + U$

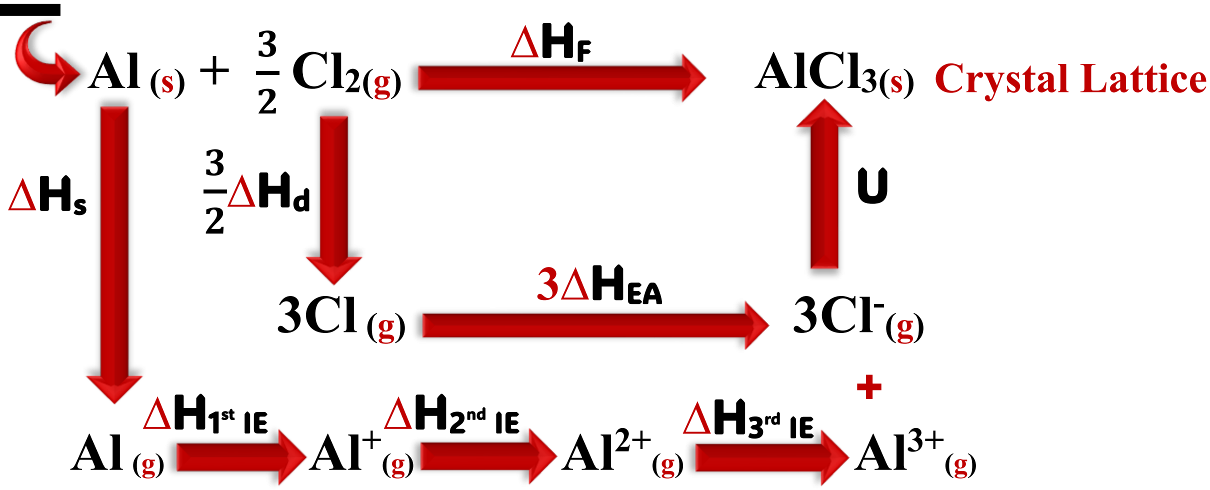
∴ $-\ 411 = 108 + (-121.1) + 500 + (-349) + U$ ∴ $U = -\ 548.9$ KJ

- **Ex 2: $MgCl_2$**



$$\boxed{\Delta H_F = \Delta H_s + \Delta H_d + \Delta H_{1^{st}IE} + \Delta H_{2^{nd}\,IE} + 2\Delta H_{EA} + U}$$

- **Ex 3: $AlCl_3$**



$$\boxed{\Delta H_F = \Delta H_s + {}^3/_2\,\Delta H_d + \Delta H_{1^{st}IE} + \Delta H_{2^{nd}IE} + \Delta H_{3^{rd}\,IE} + 3\,\Delta H_{EA} + U}$$

♦ Properties of Ionic Compounds

1) Solid, Brittle with High melting point ⟶ (Due to strength of ionic bond) نتيجة قوه الرابطة الأيونية

2) High Electrical conductivity in solution or molten state Not solid state.

3) Soluble in Polar solvent (H_2O and NH_3) and insoluble in non-polar solvent (benzene).

4) Usually organized into order lattice of atoms

Quiz

♦ Which of the following have highest lattice Energy

A) MgO or $MgCl_2$ B) NaCl or NaF

♦ Which of the following is common ion for :

A) $_{15}P$ is (P^-, P^{2-}, P^{2+}, P^{3-})

B) $_{20}Ca$ is (Ca^+, Ca^{2-}, Ca^{2+}, Ca^-)

Chapter Three

Part (2)

- ✓ **Covalent Bond**

  (Lewis Structure)

- ✓ **Coordination Bond**

2) Covalent Bond

Sharing of two electrons (One pair of e's) between two non-metals atoms

♦ Ex 1: H-H

H· ·H → H:H Or H—H

$\underbrace{\text{H:H}}$ Covalent Bond (Two Sharing e's)

(Two bonding e's)

❖ Cl_2 → Cl:Cl Or Cl—Cl

❖ HCl → H:Cl Or H—Cl

♦ Lewis Structure and Octet Rule

♦ Ex 1: $_1H$ ($1S^1$)

H· + ·H → H:H Or H—H

✔ Each H- atom becomes Surrounded by Two electrons (Isoelectronic with Helium) ∴ More stable

♦ Ex 2: $_9F$ ($1S^1$ $2s^2$ $2p^5$)

∴Valence e's = 7

:F· + ·F: → :F:F: or :F—F:

✔ Each F- atom becomes Surrounded by eight electrons (Isoelectronic with Neon) ∴ More stable

Important Notes

♦ Lewis Structure of HF

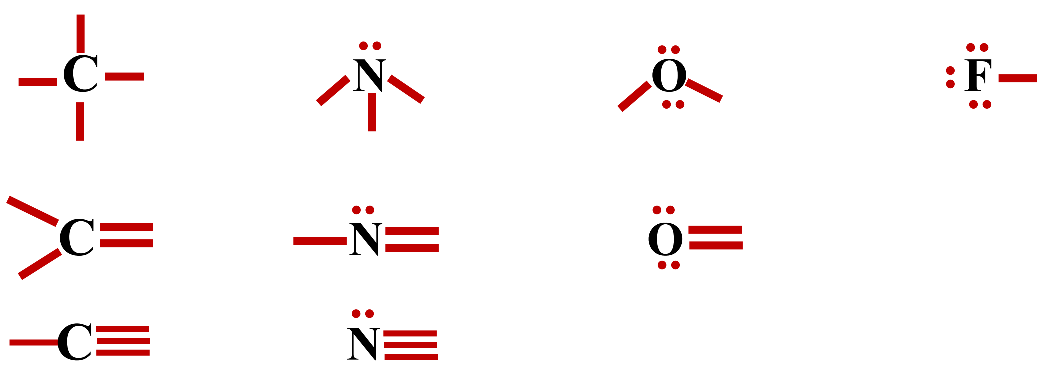
$$H\cdot + \cdot\ddot{\underset{..}{F}}: \rightarrow H—\ddot{\underset{..}{F}}:$$

Unshared e's
= Non-bonding e's
= Lone Pair of e's
∴ 3 lone Pairs

Shared e's
= bonding e's
= Covalent bond

Gp 4	Gp 5	Gp 6	Gp 7
C	N	O	F
4 bonds	3 bonds	2 bonds	1 bond
No lone pair	1 lone pair	2 lone pair	3 lone pair

Or



Or

EXAMPLE

♦ Lewis Structure of N_2 and O_2

$:N\equiv N:$ And $\ddot{\underset{..}{O}}=\ddot{\underset{..}{O}}$

Covalent Bond

Polar Covalent

Between two different atoms

(Different EN)

Ex: $^{+\delta}H \longrightarrow F^{-\delta}$, $^{+\delta}C \longrightarrow F^{-\delta}$

$^{+\delta}C=O^{-\delta}$

Non-Polar Covalent

Between Similar atoms

(Same EN)

Ex: F— F, H—H, O=O

Important Notes

1 ✔ As the difference in EN between two atoms increases bond Polarity increases

Ex Bond Polarity of : 1] C–F > C–Cl > C–Br > C–I

2] C–F > C–O > C–N > C–H

Order of EN :

حفظ

F > O > N > Cl > S > Br > I > C > H

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2 **Inert gases (Nobel gases):** $_{2}He$, $_{10}Ne$, $_{18}Ar$, $_{36}Kr$, $_{54}Xe$

3

Covalent Bond

σ-Bond | π-Bond

Covalent Bond

Single Bond	Double Bond	Trible Bond
1 σ-Bond	1 σ-Bond +1π-Bond	1 σ-Bond +2π-Bond
Cl–Cl	O=O	N≡N

∴ Bond length is Single Bond > Double Bond > Trible Bond

∴ Bond strength is Single Bond < Double Bond < Trible Bond

♦ Properties of Covalent Compounds

1) Low m.p
2) Not conduct Electricity (insulator)
3) Soluble in non-polar solvent (benzene)
4) May be Solid, Liquid or Gas

Steps for Drawing Lewis Structure

♦ **Ex: Lewis Structure of Water (H_2O)**

Steps

1 Determine the total No of Valence e's

♦ $_8O \rightarrow 1S^2\ 2S^2\ 2P^4$ (∴ Valence e's = 6 ∴ gp = 6)

♦ $_1H \rightarrow 1S^1 \rightarrow 1x2_{atoms}$ (∴ Valence e's = 2 ∴ gp = 1)

∴ Total Valence e's = 8

2 Draw a skeleton structure

♦ Put the least Electron Negativity atom in the center except H always in terminal

H–O–H

3 Determine number of unshared e's (Available)

♦ Available e's = Total Valence e's – Shared e's (bonded)

For H_2O

♦ Available e's = 8 – 4 = 4

✔ 2 bond O–H (each bond = 2 e's)

4 Determine number of e's required for each atom to obey octet Rule

♦ For H= zero ♦For any atom = 8- bonded e's

♦ H_2O: For H= zero ♦For O = 8- 4= 4 ✔

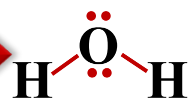
Then

♣ There are 3 cases depend on Step (3) and Step (4)

I ♦ **If Number of available e's (step 3) = number of e's needed for octet rule (step 4)**

∴ **Added Available e's as lone pair**

♦ **For H_2O: Step 3 = Step 4**



Lewis structure for H_2O

II ♦ **If (step 3) < (step 4) Only by 2 e's**

∴ **One single bond changed into One double bond**

III ♦ **If (step 3) < (step 4) Only by 4 e's**

∴ **One single bond changed into One Triple bond**

∴ **Two single bond changed into Two double bond**

Important Note

♦ **For molecules of different atoms**

∴ **Put least EN atom in center**

♦ *Ex* : CNO → N–C–O

least EN atom

1) Ammonia NH_3

1 Valence e's

♦ $_7N : 1S^2\ 2S^2\ 2P^3$ (∴ Valence e's = 5 ∴ gp = 5)

♦ $_1H : 1S^1$ (∴ 1x 3 H atoms = 3)

∴ Total Valence e's = 8

2 skeleton structure

```
H–N–H
  |
  H
```

3 Available e's = Valence e's – Shared e's (bonded)

= 8 – 6= 2 e's

4 e's required for Octet

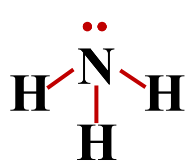
♦ For H= zero

♦ For N = 8- bonded e's

= 8 – 6 = 2

∴ <u>№ of available e's (step 3) = № of e's needed for octet rule (step 4)</u>

∴ <u>Added Available e's as lone pair on central atom</u>



<u>Lewis structure for NH_3</u>

2) Ammonium ion NH_4^+

♦ +Ve charge decrease valence e's by One

1 Valence e's

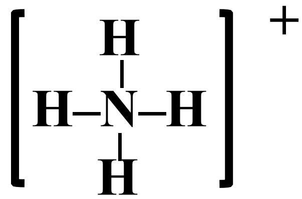
♦ $_7N : 1S^2\ 2S^2\ 2P^3$ (∴ Valence e's = 5 ∴ gp = 5)

♦ $_1H : 1S^1$ (∴ 1x 4 H atoms = 4)

♦ +Ve Charge (decrease by 1 → -1)

∴ Total Valence e's = 8

2 skeleton structure



3 Available e's = Valence e's – Shared e's (bonded)

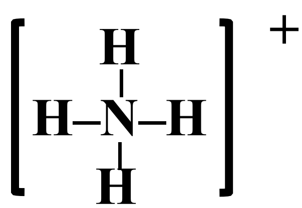
= 8 – 8= 0

4 e's required for Octet

♦ For H= zero

♦ For N = 8- bonded e's

= 8 – 8 = 0



Lewis structure for NH_4^+

3) Hydroxide ion OH^-

♦ -Ve charge increase valence e's by One

1 Valence e's

♦ $_8O : 1S^2\ 2S^2\ 2P^4$ (∴ Valence e's = 6 ∴ gp = 6)
♦ $_1H : 1S^1$ (∴ Valence e's = 1)
♦ -Ve Charge (increase by 1 → +1)

∴ Total Valence e's = 8

2 skeleton structure

$[O–H]^-$

3 Available e's

Available e's = Valence e's – Shared e's (bonded)
= 8 – 2 = 6

4 e's required for Octet

♦ For H = zero
♦ For O = 8- bonded e's
= 8 – 2 = 6

∴ If № of available e's = № of e's needed for octet rule

∴ Added Available e's as lone pair on central atom

$[:\ddot{\underset{..}{O}}–H]^-$

Lewis structure for OH^-

4) hypochlorite OCl^-

- -Ve charge increase valence e's by One

1 Valence e's

- $_8O : 1S^2\ 2S^2\ 2P^4$ (∴ Valence e's = 6)
- $_{17}Cl : 1S^2\ 2S^2\ 2P^4\ 3S^2\ 3P^5$ (∴ Valence e's = 7)
- -Ve Charge (increase by 1 → +1)

∴ Total e's = 14

2 skeleton structure $[Cl–O]^-$

3 Available e's = Valence e's – Shared e's (bonded)

= 14 – 2 = 12

4 e's required for Octet

- For Cl = 8 - 2 = 6
- For O = 8 - 2 = 6

12 e's

∴ <u>№ of available e's = № of e's needed for octet rule</u>

∴ <u>Added Available e's as lone pair</u>

$[:\ddot{\underset{..}{O}}–\ddot{\underset{..}{Cl}}:]^-$

<u>Lewis structure for OCl^-</u>

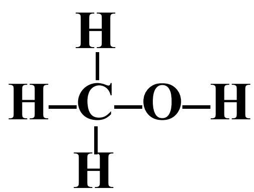
5) Methyl alcohol CH_3OH (Methanol)

1 Valence e's

- $_6C : 1S^2\ 2S^2\ 2P^2$ (∴ Valence e's = 4 ∴ gp = 4)
- $_8O : 1S^2\ 2S^2\ 2P^4$ (∴ Valence e's = 6 ∴ gp = 6)
- $_1H : 1S^1$ (∴ 1x 4 H atoms = 4)

∴ Total Valence e's = 14

2 skeleton structure



3 Available e's

Available e's = Valence e's – Shared e's (bonded)
= 14 – 10 = 4

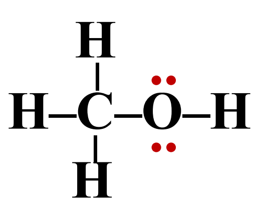
4 e's required for Octet

- For H= zero
- For O = 8- 4 = 4
- For C = 8- 8 = 0

4 e's

∴ <u>№ of available e's = № of e's needed for octet rule</u>

∴ <u>Added Available e's as lone pair</u>



Lewis structure
for CH_3OH

6) Sulphur oxide SO_2

1 Valence e's

- $_{16}S : 1S^2\ 2S^2\ 2P^4\ 3S^2\ 3P^4$ (∴ Valence e's = 6)
- $_{8}O : 1S^2\ 2S^2\ 2P^4$ (∴ Valence e's = 6 x 2)

∴ Total e's = 18

2 skeleton structure O–S–O

Less EN atom

3 Available e's = Valence e's – Shared e's (bonded)
= 18 – 4 = 14

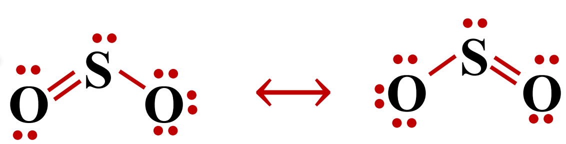
4 e's required for Octet

- For S = 8 – 4 = 4
- For O = 8 - 2 = 6 x 2

16 e's

∴ Step 4 > Step 3 by 2 e's

∴ One single bond changed into One double bond



2 Resonance Structure

- **Resonance**

The presence of molecules in more one than Lewis structure

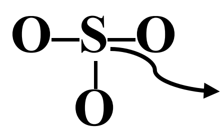
7) Sulphur Trioxide SO_3

1 Valence e's

- $_{16}S : 1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^4$ (∴ Valence e's = 6)
- $_{8}O : 1S^2\ 2S^2\ 2P^4$ (∴ Valence e's = 6 x 3)

∴ Total e's = 24

2 skeleton structure



Less EN atom

3 Available e's = Valence e's – Shared e's (bonded)

= 24 – 6 = 18

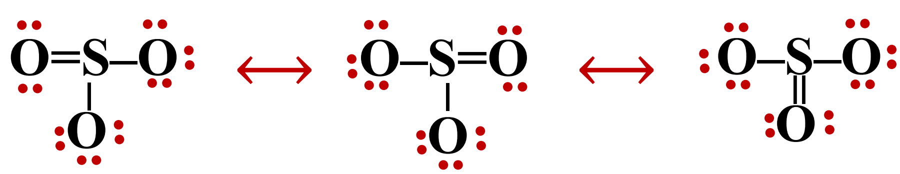
4 e's required for Octet

- For S = 8 – 6 = 2
- For O = 8 - 2 = 6 x 3 atoms

20 e's

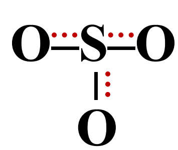
∴ Step 4 > Step 3 by 2 e's

∴ One single bond changed into One double bond



3 Resonance Structure

∴ Actual structure is the average of the 3 resonances



Resonance hybrid structure

∴ (Electrons are de-localized moves around the entire molecule not localized between two atoms)

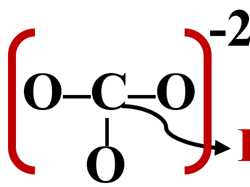
8) Carbonate CO_3^{2-}

1 Valence e's

- $_6C$: $1S^2\ 2S^2\ 2P^2$ (∴ Valence e's = 4)
- $_8O$: $1S^2\ 2S^2\ 2P^4$ (∴ Valence e's = 6 x 3)
- 2 -Ve Charge (increase by 2 → +2)

∴ Total e's = 24

2 skeleton structure



Less EN atom

3 Available e's = Valence e's – Shared e's (bonded)

= 24 – 6 = 18

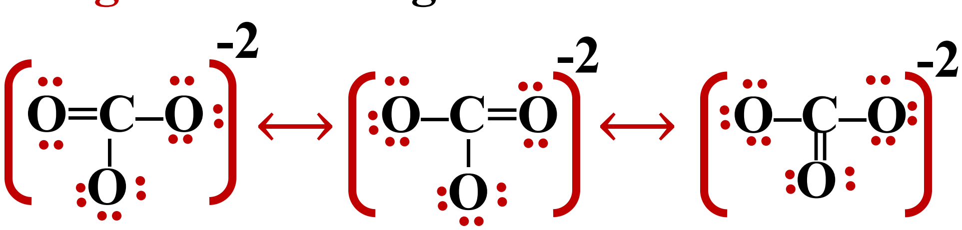
4 e's required for Octet

- For C = 8 – 6 = 2
- For O = 8 - 2 = 6 x 3

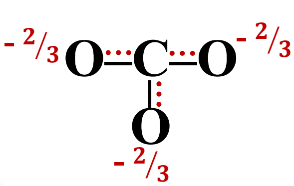
20 e's

∴ Step 4 > Step 3 by 2 e's

∴ One single bond changed into One double bond



3 Resonance Structure



Resonance hybrid structure

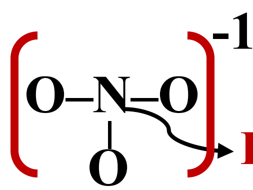
9) Nitrate NO_3^-

1 Valence e's

- $_7N$: $1S^2\ 2S^2\ 2P^2$ (∴ Valence e's = 5)
- $_8O$: $1S^2\ 2S^2\ 2P^4$ (∴ Valence e's = 6 x 3)
- -Ve Charge (increase by 1 → +1)

∴ Total e's = 24

2 skeleton structure



Less EN atom

3 Available e's = Valence e's – Shared e's (bonded)
= 24 – 6 = 18

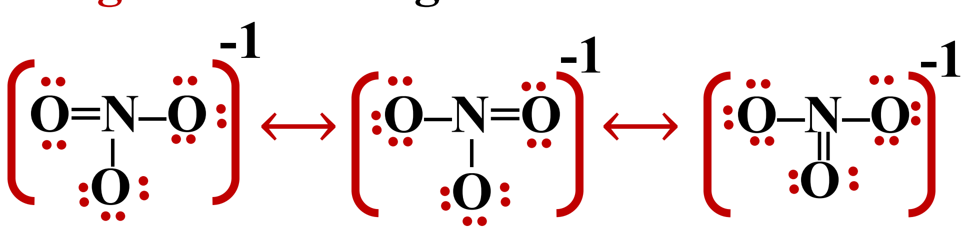
4 e's required for Octet

- For N = 8 – 6 = 2
- For O = 8 - 2 = 6 x 3

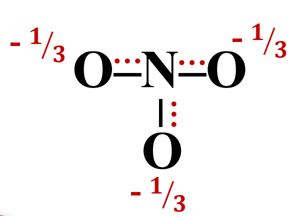
20 e's

∴ Step 4 > Step 3 by 2 e's

∴ One single bond changed into One double bond



3 Resonance Structure



Resonance hybrid structure

10) Nitrogen molecule N_2

1 Valence e's ♦ $_7N : 1S^2\ 2S^2\ 2P^3$ (∴ Valence e's = 5x2)

∴ Total Valence e's = 10

2 skeleton structure N–N

3 Available e's = Valence e's – Shared e's (bonded)

= 10 – 2= 8 e's

4 e's required for Octet

♦For N = 8- bonded e's

= (8 – 2) x 2 $_{atoms}$ = 12

∴ Step 4 > Step 3 by 4 e's

∴ One single bond changed into One Trible bond

$\ddot{N}\equiv\ddot{N}$

Lewis structure for N_2

Formal Charge C_F

It is possible to draw different Lewis structures for molecule with different arrangement of atoms

Ex : Molecule AB_2

Possible Lewis structure may be

A-B-B **Or** B-A-B

But:

♦ **Only one lewis structure is stable and exist**

∴ **Stability of structure is determined from Formal Charge (C_f)**

Formal charge of any atom is given by :

$$C_f = \underbrace{\text{Valence e's of atom}}_{\text{gp № of atom}} - (\text{unshared e's} + \text{№ of Bonds})$$

Where :

A] C_f for all atoms are **zero** or **close to zero**

B] The **-Ve** C_f is located on the **more EN** atom

1) Carbon dioxide CO_2

1 Valence e's

- $_6C : 1S^2\ 2S^2\ 2P^2$ ($\therefore$ Valence e's = 4)
- $_8O : 1S^2\ 2S^2\ 2P^4$ ($\therefore$ Valence e's = 6 x 2)

$\therefore$ Total e's = 16

2 skeleton structure O–C–O

C → Less EN atom

3 Available e's = 16 – 4 = 12

4 e's required for Octet

- For C = 8 – 4 = 4
- For O = 8 - 2 = 6 x 2

16 e's

$\therefore$ Step 4 > Step 3 by 4 e's

One single → One Triple

$\ddot{O}\equiv C-\ddot{\underset{\cdot\cdot}{O}}:$

- $C_{F\ O_{left}} = 6 - (2+3) = +1$
- $C_{F\ C} = 4 - (0+4) = 0$
- $C_{F\ O_{right}} = 6 - (6+1) = -1$

$\therefore$ **Less stable**

Two single → Two Double

$\underset{\cdot\cdot}{\ddot{O}}=C=\underset{\cdot\cdot}{\ddot{O}}$

- $C_{F\ O_{left}} = 6 - (4+2) = 0$
- $C_{F\ C} = 4 - (0+4) = 0$
- $C_{F\ O_{right}} = 6 - (4+2) = 0$

$\therefore$ **More stable**

Lewis structure for CO_2

2) Draw NOC^- and predict more stable structure

1 Valence e's

- $_7N$: $1S^2\ 2S^2\ 2P^3$ (∴ Valence e's = 5)
- $_6C$: $1S^2\ 2S^2\ 2P^2$ (∴ Valence e's = 4)
- $_8O$: $1S^2\ 2S^2\ 2P^4$ (∴ Valence e's = 6)
- -Ve Charge (increase by 1 → +1)

∴ Total e's = 16

2 skeleton structure $[N-C-O]^-$

Less EN atom (C)

3 Available e's = 16 – 4 = 12

4 e's required for Octet

- For C = 8 – 4 = 4
- For O = 8 - 2 = 6
- For N = 8 - 2 = 6

16 e's

∴ Step 4 > Step 3 by 4 e's

One single → One Triple | **Two single → Two Double**

(I) $[\ddot{N}\equiv C-\ddot{\underset{\cdot\cdot}{O}}:]^-$

- $C_{F\,N}$ = 5 - (2+3) = 0
- $C_{F\,C}$ = 4 - (0+4) = 0
- $C_{F\,O}$ = 6 - (6+1) = -1

✓ ∴ **More stable**

$[:\ddot{\underset{\cdot\cdot}{N}}-C\equiv \ddot{O}]^-$ (II)

- $C_{F\,N}$ = 5 - (6+1) = -2
- $C_{F\,C}$ = 4 - (0+4) = 0
- $C_{F\,O}$ = 6 - (2+3) = 1

∴ **Less stable**

$[\ddot{\underset{\cdot\cdot}{N}}=C=\ddot{\underset{\cdot\cdot}{O}}]^-$ (III)

- $C_{F\,N}$ = 5 - (4+2) = -1
- $C_{F\,C}$ = 4 - (0+4) = 0
- $C_{F\,O}$ = 6 - (4+2) = 0

∴ **Less stable**

Due to -1 located on more EN atom (O) so it is most stable structure ∴ I> III> II

3) Draw N_2O and predict more stable structure

1 Valence e's

- $_7N : 1S^2\ 2S^2\ 2P^3$ (∴ Valence e's = 5 x2)
- $_8O : 1S^2\ 2S^2\ 2P^4$ (∴ Valence e's = 6)

∴ Total e's = 16

2 skeleton structure N–N–O → Less EN atom

3 Available e's = 16 – 4 = 12

4 e's required for Octet

- For N = 8 – 4 = 4
- For O = 8 - 2 = 6
- For N = 8 - 2 = 6

16 e's

∴ Step 4 > Step 3 by 4 e's

One single → One Triple | Two single → Two Double

(I) $\ddot{N}\equiv N-\ddot{\underset{..}{O}}:$

- $C_{F\,N}$ = 5 - (2+3) = 0
- $C_{F\,N}$ = 5 - (0+4) = 1
- $C_{F\,O}$ = 6 - (6+1) = -1

✓ ∴ More stable

(II) $:\underset{..}{\ddot{N}}-N\equiv\ddot{O}$

- $C_{F\,N}$ = 5 - (6+1) = -2
- $C_{F\,N}$ = 4 - (0+4) = 1
- $C_{F\,O}$ = 6 - (2+3) = 1

∴ Less stable

(III) $\underset{..}{\ddot{N}}=N=\underset{..}{\ddot{O}}$

- $C_{F\,N}$ = 5 - (4+2) = -1
- $C_{F\,N}$ = 4 - (0+4) = 1
- $C_{F\,O}$ = 6 - (4+2) = 0

∴ Less stable

Due to -1 located on more EN atom (O) so it is most stable structure (EN of O > N) ∴ I> III> II

Exception of Octet Rule

1) Molecules contain odd number of valence e's and don't obey octet Rule

Ex NO

1 Valence e's

♦ $_7N$: $1S^2\ 2S^2\ 2P^3$ (∴ Valence e's = 5)

♦ $_8O$: $1S^2\ 2S^2\ 2P^4$ (∴ Valence e's = 6)

∴ Total e's = 11 (odd فردى)

2 skeleton structure N–O

3 Available e's = Valence e's – Shared e's (bonded)

= 11 – 2 = 9

4 e's required for Octet

♦ For N = 8 – 2 = 6

♦ For O = 8 - 2 = 6

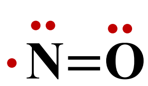
12 e's

∴ Step 4 > Step 3 by 3 e's

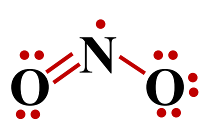
∴ One single bond changed into One double bond

∴ Distribute remaining available e's as unshared e's

Not obey the octet rule ←



Ex 2 NO_2 ∴ Valence e's = 17 (odd فردى)



حفظ

2) Molecules in which central atom surrounded by only two or three pair of e's rather than four pairs

Ex 1 $BeCl_2$, BeF_2 or BeX_2

Lewis structure according to the octet rule by applying steps

$\overset{+1}{Cl}=\overset{-2}{Be}=\overset{+1}{Cl}$

- $C_{F\ Cl} = +1$
- $C_{F\ Be} = -2$

∴ **Less stable does not exist**

True Lewis Structure (Exist)

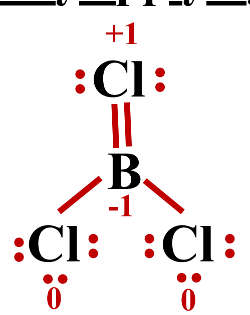
$:\overset{0}{Cl}-\overset{0}{Be}-\overset{0}{Cl}:$

- $C_{F\ Cl} = 0$
- $C_{F\ Be} = 0$

∴ **More stable exist**

Ex 2 BCl_3, BF_3 or BX_3

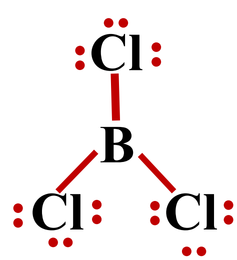
Lewis structure according to the octet rule by applying steps



- $C_{F\ Cl} = +1, 0, 0$
- $C_{F\ B} = -1$

∴ **Less stable does not exist**

True Lewis Structure (Exist)



- $C_{F\ Cl} = 0$
- $C_{F\ B} = 0$

∴ **More stable exist**

3) Expanded Octet

Molecules in central atoms are surrounded by more than four Pairs of e's

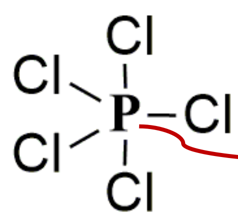
Ex 1 PCl_5 (Phosphorus Penta chloride)

1 Valence e's

- $_{15}P$: gp 5 (∴ Valence e's = 5)
- $_{17}Cl$: gp 7 (∴ Valence e's = 7 x 5)

∴ Total Valence e's = 40

2 skeleton structure



Expanded octet (P) surrounded by 10 e's

3 Available e's = 40 – 10 = 30

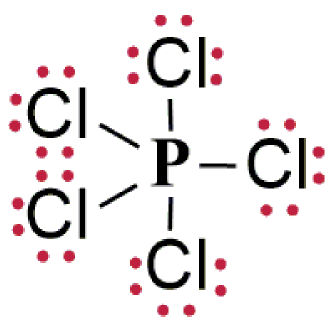
4 e's required for Octet

- For P = zero
- For Cl = 8 - 2 = 6 x 5

30 e's

فى حاله Expanded Octet نقوم بتوزيع الالكترونات المتاحه ك lone pair على الذرات الطرفيه

لكى تصل الى Octet و لو فى زياده من الالكترونات نضعها على ذره مركزيه ك lone pair



Lewis structure for PCl_5

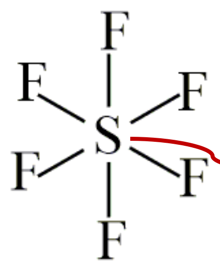
Ex 2 SF_6 (Sulphur hexafluoride)

1 Valence e's

- $_{16}S$: gp 6 (∴ Valence e's = 6)
- $_{9}F$: gp 7 (∴ Valence e's = 7 x 6)

∴ Total Valence e's = 48

2 skeleton structure



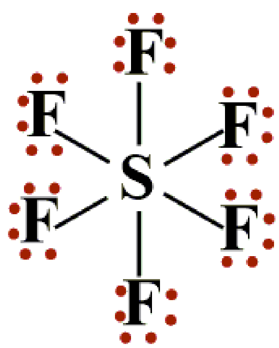
Expanded octet (S) surrounded by 12 e's

3 Available e's = 48 – 12 = 36

4 e's required for Octet

- For S = zero
- For F = 8 - 2 = 6 x 6

36 e's



Lewis structure for SF_6

Ex 3 SF_4 (Sulphur Tetrafluoride)

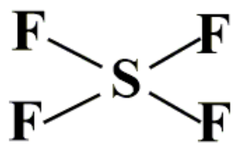
1 Valence e's

- $_{16}S$: gp 6
- $_{9}F$: gp 7

(∴ Valence e's = 6)
(∴ Valence e's = 7 x 4)

∴ Total Valence e's = 34

2 skeleton structure



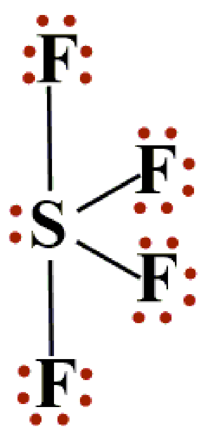
3 Available e's = 34 – 8 =26

4 e's required for Octet

- For S = zero
- For F = 8 - 2 = 6 x 4

24 e's

∴ Step 3 > Step 4 (Expanded Octet)



Lewis structure for SF_4

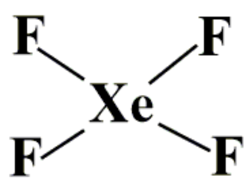
Ex 4 XeF_4

1 Valence e's

- $_{54}Xe$: gp 8 (∴ Valence e's = 8 Nobel gas)
- $_{9}F$: gp 7 (∴ Valence e's = 7 x 4)

∴ Total Valence e's = 36

2 skeleton structure



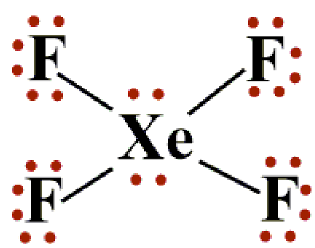
3 Available e's = 36 – 8 = 28

4 e's required for Octet

- For Xe = zero
- For F = 8 - 2 = 6 x 4

24 e's

∴ Step 3 > Step 4 (Expanded Octet)



Lewis structure for XeF_4

Ex 5 $XeCl_2$

1 Valence e's

- $_{54}Xe$: gp 8 (∴ Valence e's = 8 Nobel gas)
- $_{17}Cl$: gp 7 (∴ Valence e's = 7 x 2)

∴ Total Valence e's = 22

2 skeleton structure Cl–Xe–Cl

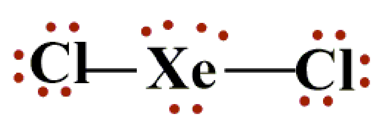
3 Available e's = 22 – 4 = 18

4 e's required for Octet

- For Xe = 4
- For Cl = 6 x 2

16 e's

∴ Step 3 > Step 4 (Expanded Octet)



Lewis structure for $XeCl_2$

Ex 6 ClF_3

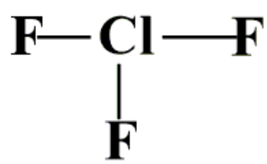
1 Valence e's

♦ $_{17}Cl$: gp 7
♦ $_{9}F$: gp 7

(∴ Valence e's = 7)
(∴ Valence e's = 7 x 3)

∴ Total Valence e's = 28

2 skeleton structure



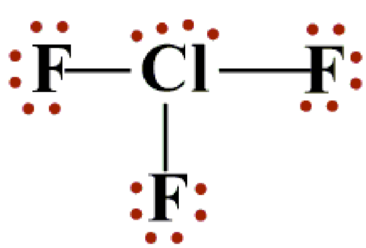
3 Available e's = 28 – 6 =22

4 e's required for Octet

♦ For Cl = 2
♦ For F = 6 x 3

20 e's

∴ Step 3 > Step 4 (Expanded Octet)



Lewis structure for ClF_3

♣ Note Be

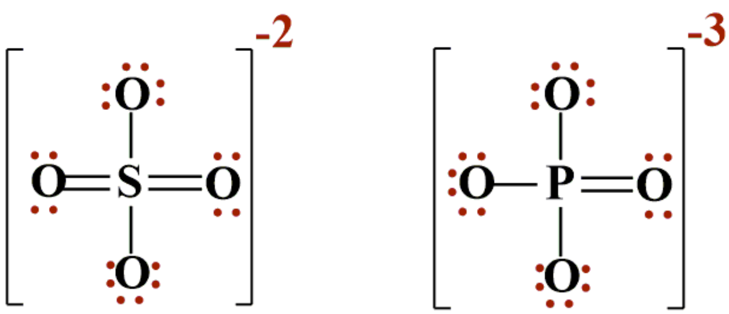
1] Any atom containing empty d-orbitals can become central atom in an expanded octet

Ex: Phosphorus can be formed PCl_5 while nitrogen can't form NCl_5

Because
- ♦ ${}_7N$: $1S^2\ 2S^2\ 2P^3$
- ♦ ${}_{15}P$: $1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^3\ 3d^0$

∴ P-atom contains empty 3d-orbitals that accept extra pairs of e's and form expanded octet PCl_5, while N-atom not contain 3d-orbitals

2] Structure of SO_4^{2-} and PO_4^{3-} (حفظ)



3] All Halogen can be Put as a central atom, Except F always as a terminal atom (has the highest EN)

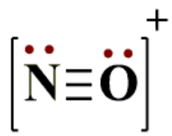
Quiz

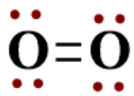
♦ Draw Lewis structure for the following :

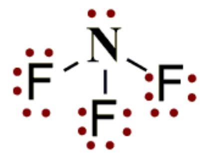
A) NO^+ B) O_2 C) NF_3 D) CH_4

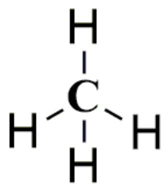
E) HCN F) H_3O^+ G) O_3 H) $COCl_2$ I) HCO_2^-

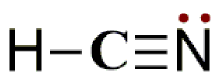
Answer

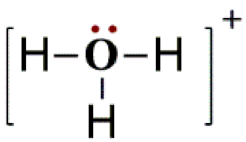
A) 

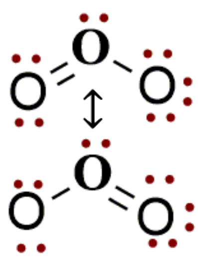
B) 

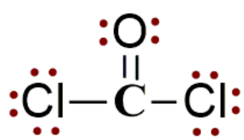
C) 

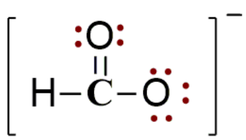
D) 

E) 

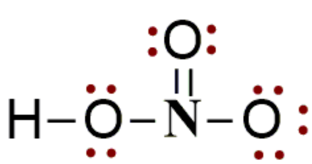
F) 

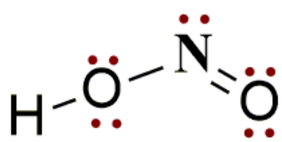
G) 

H) 

I) 

Note (حفظ)

1- HNO_3 

2- HNO_2 

3- Coordination Bond

Formed between metal ion (vacant orbitals) with molecule or ions (rich electrons) called ligand.

∴ Metal ions: electron acceptor

♦ is called Lewis Acid

∴ Ligand: electron donor

♦ is called Lewis Base

e's transfer

M←Ligand (coordinate bond)

Lewis Acid

Is an electron acceptor or electrophile species having vacant orbitals to accept lone pairs of electrons. (e's Deficient)

Ex: Fe^{2+}, Cu^{2+}, Fe^{3+}, Co^{2+}, Co^{3+},........

Lewis Base

Is an electron donor species having at least one lone pair of electrons which can be donated to an atom acting electron acceptor and ligand can be neutral or negative charged

Ex: Neutral molecule (NH_3, H_2O) or ion (Cl^-, CN^-, SO_4^{2-})

✓ **Note Be**

Coordination number is the Number of coordinate bonds depend on № of donor atoms (centers) in ligand surrounded the metal ion.

Types Of ligands

According to № of donor atoms

Monodentate

- One donor atom

Ex: H_2O, NH_3, Cl^-

Bidentate

- Two donor atoms

Ex: Ethylene diamine

$H_2\ddot{N}-CH_2-CH_2-\ddot{N}H_2$

polydentate

- more than two donor atoms

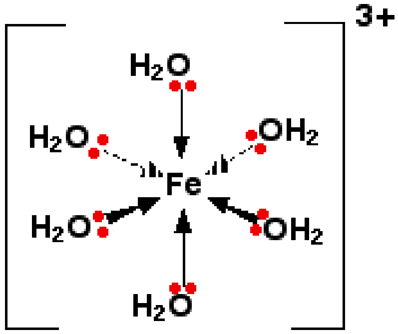
Ex: EDTA

(6 donor atoms)

✔ **Note: Bidentate and Polydentate is called chelating agent**

- **What is the coordination number in the following complex:**

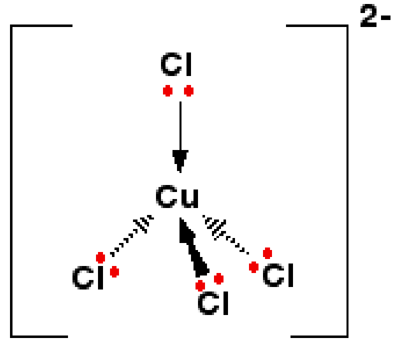
1]



∴ 6 donor atoms (O)

∴ Coordination number = 6

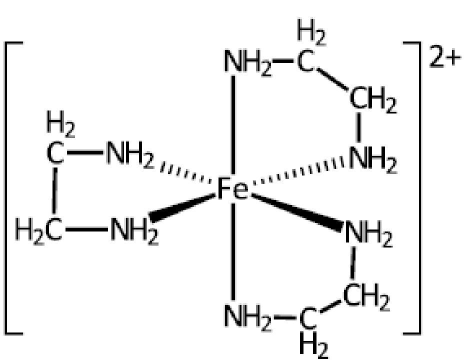
2]



∴ 4 donor atoms (Cl)

∴ Coordination number = 4

3]

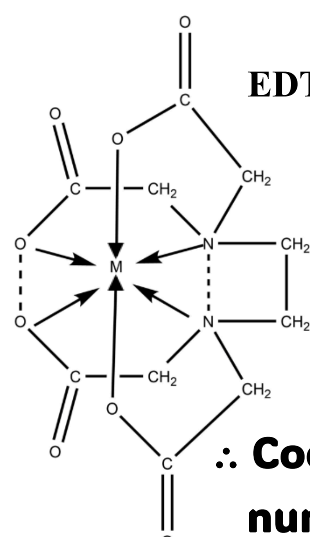


∴ 6 donor atoms (N)

∴ Coordination number = 6

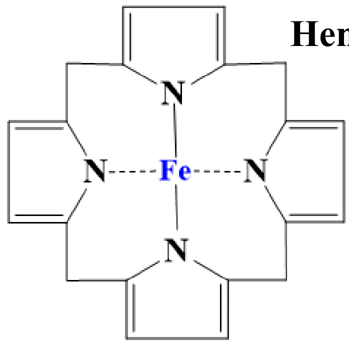
4]

EDTA



∴ Coordination number = 6

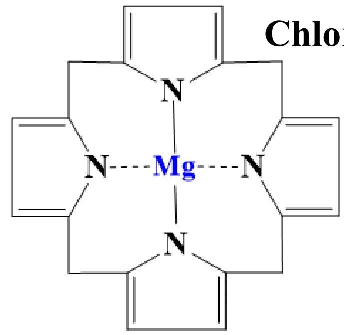
5] Hemoglobin



∴ **4 donor atoms (N)**

∴ **Coordination number = 4**

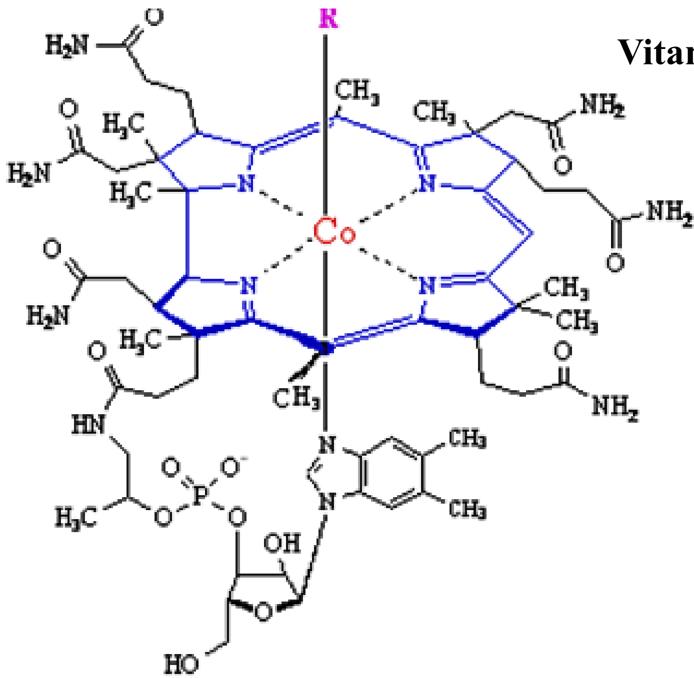
6] Chlorophyl



∴ **4 donor atoms (N)**

∴ **Coordination number = 4**

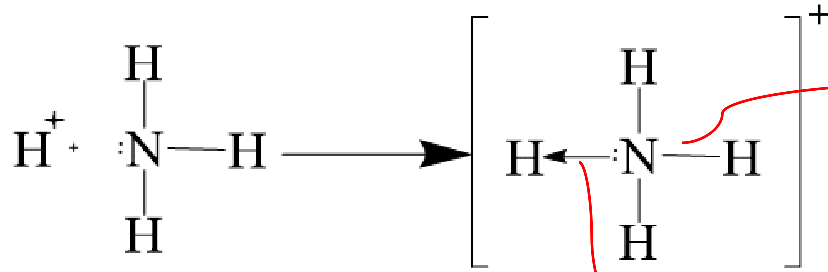
6] Vitamin B12



∴ **Coordination number = 6**

♦ **What are the types of bons exist in the following molecules:**

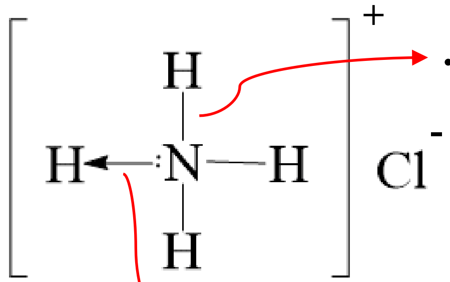
1]



∴ 3 covalent bonds

∴ 1 coordinate bond

2] NH_4Cl



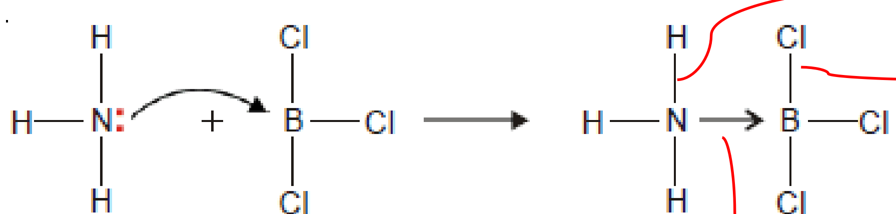
∴ 3 covalent bonds

∴ 1 ionic bond

Between NH_4^+ and Cl^-

∴ 1 coordinate bond

3] $BCl_3 + NH_3$ ($H_3N{:}\rightarrow BCl_3$)



∴ Covalent bond

∴ Covalent bond

∴Coordinate bond

Chapter Three

Part (3)

- ✓ VSEPR

VSEPR Model — Molecular Geometry

(Valence Shell Electron Pair Repulsion Model)

(VSEPR)

This model is used to determine the geometry of molecule.

Possible arrangement of atoms around central atom

♦ Order of Repulsion between valence e's pair:

Repulsion Between lone pair-lone pair (.. ..) > **Repulsion Between lone pair-bond pair** (.. –) > **Repulsion Between bond pair-bond pair** (– –)

Due to repulsion: orbitals containing that electron-pair are oriented in a certain geometry

نتيجه التنافر الذى يحدث على الذره المركزيه يبدء المركب فى تغير شكله الفراغى

Applying the VSEPR Model to Predict Geometry

♦ Steps:

1] Draw Lewis Structure

2] Represent → ♦ Central Atom = A

♦ Terminal Atom = B or (X)

♦ Lone Pair = E (on central atom only)

♦ **Ex:**

♣ $CO_2 = AB_2$ ♣ $CH_4 = AB_4$ ♣ $\ddot{N}H_3 = AB_3E$

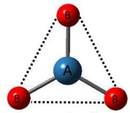
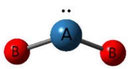
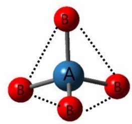
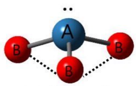
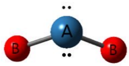
♦ Then:

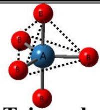
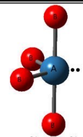
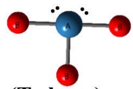
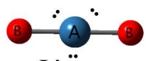
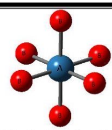
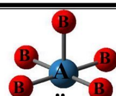
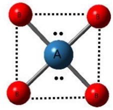
3] Use The following table on next page حفظ

Note Be :

A] № of (B+E) = Electron domain (electron gp) (effective pairs)

B] Geometry = Electron domain Geometry (E.D.G)

Electron Domain (B+E)	Type	Hybridization	Geometry	Shape	Bond angle	Ex
2	AB_2	SP	Linear	 Linear	180^0	CO_2 BeF_2
3	AB_3	SP^2	Trigonal planar	 Trigonal planar	120^0	SO_3 BF_3 CO_3^{-2} NO_3^-
	AB_2E			 Bent shape (V-shape)	$< 120^0$	SO_2 NO_2^-
4	AB_4	SP^3	Tetrahedral	 Tetrahedral	109.5^0	CH_4 NH_4^+ SO_4^{-2}
	AB_3E			 Trigonal Pyramidal	$<109.5^0$	NH_3
	AB_2E_2			 Bent shape (V-shape)	$<<109.5^0$	H_2O

(B+E)	Type	Hybridization	Geometry	Shape	angle	Ex
5	AB_5	SP^3d	Trigonal Bipyramidal	 Trigonal Bipyramidal	180^0 120^0 90^0	PCl_5
	AB_4E			 See-Saw	$< 180^0$ $< 120^0$ $< 90^0$	SF_4
	AB_3E_2			 (T-shape)	$<< 180^0$ $<< 90^0$	ClF_3
	AB_2E_3			 Linear	$\approx 180^0$	XeF_2
6	AB_6	SP^3d^2	Octahedral	 Octahedral	180^0 90^0	SF_6 XeF_6^{2+}
	AB_5E			 Square pyramidal	180^0 90^0	$BrCl_5$ ClF_5 ICl_5
	AB_4E_2			 Square Planar	180^0 90^0	BrF_4^+ XeF_4

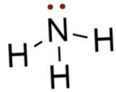
Note Be :

A] Electron Geometry →Geometry depends on № of B+E

B] Molecular Geometry → shape depends on № of B

Example :

1- NH_3

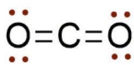
♦ Lewis Structure of Ammonia 

∴ Type AB_3E (Electron domain = 4 or effective pairs =4)

∴ sp^3 — Tetrahedral Geometry
— Trigonal Pyramidal shape

∴ Bond angle < 109.5^0

2- CO_2

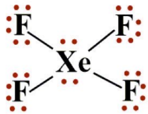
♦ Lewis Structure of CO_2 

∴ Type AB_2 (Electro domain = 2)

∴ sp — linear Geometry
— linear shape

∴ Bond angle = 180^0

3- XeF_4



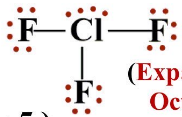
(Expanded Octet)

- ♦ Lewis Structure of XeF_4
  - ∴ Type AB_4E_4 (Electro domain = 6)

∴ sp^3d^2 hybridization
- Octahedral Geometry
- Square planar shape

∴ Bond angle = 90^0 and 180^0

4- ClF_3



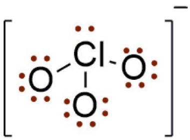
(Expanded Octet)

- ♦ Lewis Structure of ClF_3
  - ∴ Type AB_3E_2 (Electron domain = 5)

∴ sp^3d hybridization
- Trigonal bipyramidal Geometry
- T-shape

∴ Bond angle $\ll 90^0$ and $\ll 180^0$

5- ClO_3^-



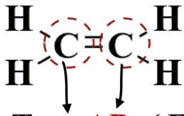
- ♦ Lewis Structure of ClO_3^-
  - ∴ Type AB_3E (Electron domain = 4)

∴ sp^3 hybridization
- Tetrahedral Geometry
- Trigonal pyramidal shape

∴ Bond angle < 109.5^0

Note Be : لو عندى ذرتين مركزيتين اخد كل ذره لوحدها كمركز لوحده

6- C_2H_4

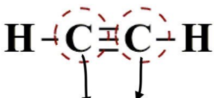


∴ Bond angle $= 120^0$

∴ Type AB_3 (Electron domain = 3)

∴ sp^2 hybridization
- Trigonal planar Geometry
- Trigonal planar shape

7- C_2H_2



∴ Bond angle $= 180^0$

∴ Type AB_2 (Electron domain = 2)

∴ sp hybridization
- Linear Geometry
- Linear shape

Quiz

♦ Predict the electron geometry and shape

A) NO_3^- B) PCl_5 C) BeH_2

D) IF_5 E) BCl_3 F) CH_4 G) AX_4E_2

♦ Order the following according to the increasing in bond angle

A) NO_3^-, NO_2^-, NH_4^+

B) H_2O, NH_3, NH_4^+

B) H_2O, BCl_3, CH_4

Chapter Three

Part (4)

- ✓ VBT

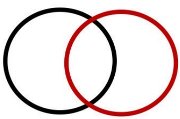
Valence Bond Theory (VBT) (Hybridization)

- **Lewis's theory does not clearly explain why chemical bonds exist**
- **Valence bond theory based on quantum mechanics which explain how bond formation**
- **The covalent bond formed due to overlapping between two atomic orbitals where each orbital contains one electron**

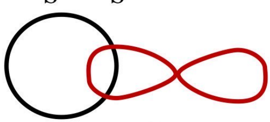
Covalent Bond

σ-Bond (Strong)

(Result from Overlapping between)



S S



S P

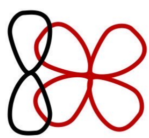
P P

Head-to-Head

π-Bond (weak)

(Result from Overlapping between)

P P



P d

Side-to-Side

Explain covalent bond formation (Hybridization)

Ex 1 $AB_2 \rightarrow BeH_2$

♦ $_4Be : 1S^2\ 2S^2\ 2P^0$ ∴ Valence shell

$2P^0$ [| |]

$2S^2$ [⇅]

e⁻-Promotion (Excitation)

[↿| |] $2P^1$

$2S^1$ [↿]

$\underbrace{\text{One s-orbital + One p-orbital}}$

Two sp-Hybrids orbitals

[↿|↿] 2sp [|] 2p

Bond Formation

[⇅|⇅] sp

2 Be-H bonds

s-orbital s-orbital

∴ H–Be–H

∴ sp – linear Geometry
– linear shape

∴ Bond angle = 180^0

∴ Effective electron pair (Electron domain gp) = 2

Ex 2 $AB_3 \rightarrow BF_3$

♦ $_5B$: $1S^2\ 2S^2\ 2P^1$ ∴ Valence shell

[↿ | |] $2P^1$

$2S^2$ [⇅]

e^--Promotion (Excitation)

[↿ | ↿ |] $2P^2$

$2S^1$ [↿]

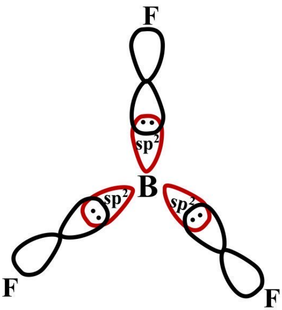
One s-orbital + Two p-orbital (Hybridization)

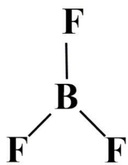
Three sp^2-Hybrids orbitals [↿ | ↿ | ↿] sp^2 [ ] 2p

Bond Formation

[⇅ | ⇅ | ⇅] sp^2

3 B-F Bonds





∴ sp^2 — Trigonal planar Geometry / Trigonal planar shape

∴ Bond angle = 120^0

∴ Effective electron group = 3

Note Be : ♦ $_7F$: $1S^2\ 2S^2\ 2P^5$

[⇅ | ⇅ | ↿] → 1 unpaired e^- exist in p form a covalent bond

Ex 3 $AB_4 \rightarrow CH_4, CCl_4$

♦ $_6C$: $1S^2\ 2S^2\ 2P^2$ ∴ Valence shell

$2P^2$: [↿ | ↿ |]

$2S^2$: [↿⇂]

e⁻-Promotion (Excitation)

$2P^3$: [↿ | ↿ | ↿]

$2S^1$: [↿]

$\underbrace{\text{One s-orbital + Three p-orbital}}$ (Hybridization)

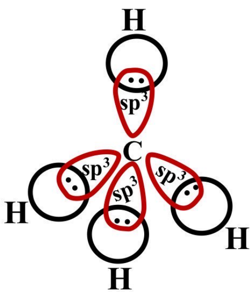
Four sp^3-Hybrids orbitals

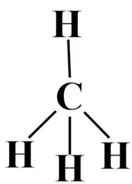
[↿ | ↿ | ↿ | ↿] Sp^3

Bond Formation

[↿⇂ | ↿⇂ | ↿⇂ | ↿⇂] Sp^3

4 C-H Bond





∴ sp^3 — Tetrahedral Geometry / Tetrahedral shape

∴ Bond angle = 109.5^0

∴ Effective electron group = 4

Ex 4 $AB_3E \rightarrow NH_3$

e^--Promotion (Excitation)

♦ $_7N$: $1S^2$ $2S^2$ $2P^3$ ∴ Valence shell

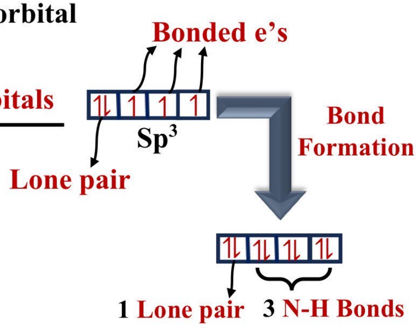
$2p^3$ ↿ ↿ ↿

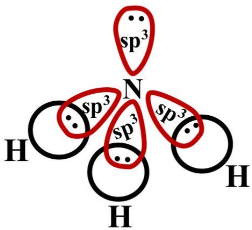
$2s^2$ ⇅

One s-orbital + Three p-orbital

(Hybridization)

Four sp^3-Hybrids orbitals





∴ sp^3 — Tetrahedral Geometry / Trigonal pyramidal shape

∴ Bond angle < 109.5^0

∴ Effective electron group = 4

Ex 5 $AB_2E_2 \rightarrow H_2O$

♦ $_8O$: $1S^2\ 2S^2\ 2P^4$ ∴ Valence shell

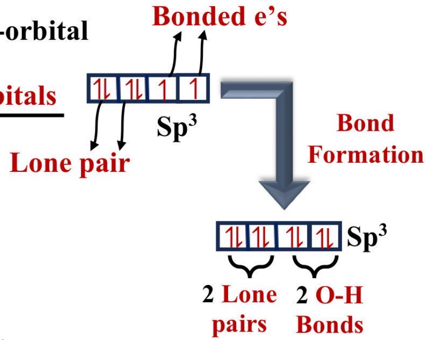
2P⁴: [⇅ | ↿ | ↿]

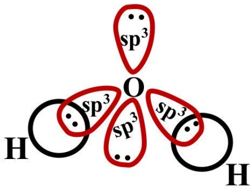
2S²: [⇅]

One s-orbital + Three p-orbital

(Hybridization)

Four sp³-Hybrids orbitals





∴ sp^3 — Tetrahedral Geometry / Bent (V-shape)

∴ Bond angle << 109.5^0

∴ Effective electron group = 4

Ex 6 $AB_5 \rightarrow PCl_5$

♦ $_{15}P$: $1S^2\ 2S^2\ 2P^6\ 3S^2\ \ 3P^3\ \ 3d^0$ ∴ Valence shell

$3p^3$ [1][1][1] [ ][ ][ ][ ][ ] $3d^0$

$3s^2$ [1⇂]

e⁻-Promotion (Excitation)

[1][ ][ ][ ][ ] $3d^1$

[1][1][1] $3p^3$

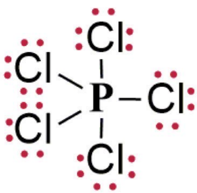
[1] $3s^1$

$\underbrace{\text{1 s-orbital + 3 p-orbital + 1 d- orbitals}}$

(Hybridization)

Five sp^3d-Hybrids orbitals [1][1][1][1][1] Sp^3d [ ][ ][ ][ ] 3d

Bond Formation



[1⇂][1⇂][1⇂][1⇂][1⇂] Sp^3d

$\underbrace{\qquad\qquad}_{\text{5 P-Cl Bonds}}$

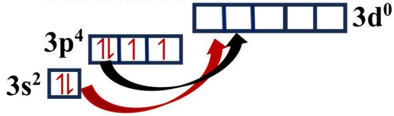
∴ sp^3d – Trigonal bipyramidal Geometry
– Trigonal bipyramidal shape

∴ Bond angle= 90^0, 120^0 and 180^0

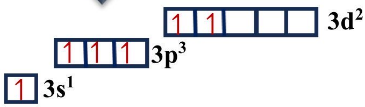
∴ Effective electron group = 5

Ex 7 $AB_6 \rightarrow SF_6$

- $_{16}S$: $1S^2\ 2S^2\ 2P^6\ 3S^2\ \ 3P^4\ \ 3d^0$ ∴ Valence shell

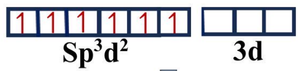


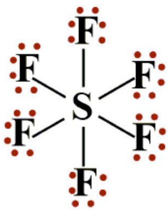
e^--Promotion (Excitation)



$$\underbrace{\text{1 s-orbital + 3 p-orbital + 2 d- orbitals}}$$

Six sp^3d^2-Hybrids orbitals





Bond Formation

6 S-F Bonds

∴ sp^3d — Octahedral Geometry / Octahedral shape

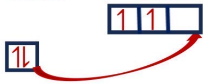
∴ Bond angle= 90^0 and 180^0

∴ Effective electron group = 6

Ex 8 Hybridization of ethylene → $CH_2=CH_2$

♦ $_6C$: $1S^2$ $2S^2$ $2P^2$

$CH_2= CH_2$ → Sp^2



e^--Promotion (Excitation)

Hybridization between One s-orbital and Two p-orbital to give three sp^2 hybrid orbitals

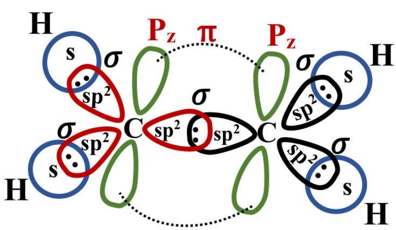
while

Third Pz orbital remains unhybrid to form π-bond

three sp^2-Hybrids orbitals

sp^2 — Unhybrid P_z

Bond Formation



sp^2 — 3 σ Bonds; 1 π Bond

♦ C-H: σ-Bond (sp^2-s)

♦ C-C: σ-Bond (sp^2- sp^2)

♦ C=C: π-Bond (p_z- p_z)

$CH_2= CH_2$

♣ Effective electron gp= 3

♣ sp^2-hybridization

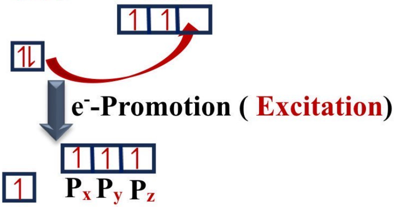
♣ Trigonal planar geometry and shape

♣ Bond angle = 120^0

Ex 9 Hybridization of Acetylene → CH≡CH

$\underbrace{CH{\equiv}CH}_{Sp}$

♦ $_6C$: $1S^2$ $2S^2$ $2P^2$

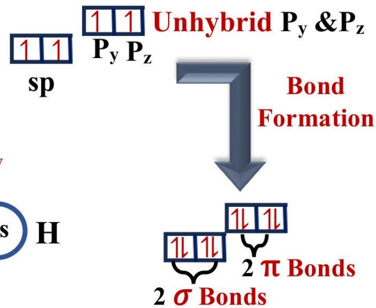


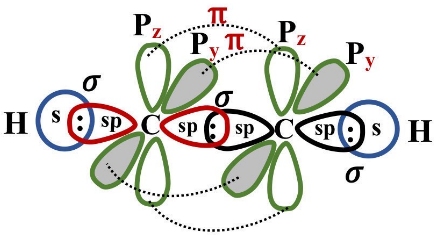
Hybridization between One s-orbital and One p-orbital to give Two sp hybrid orbitals

while

P_y and P_z orbitals remains unhybrid to form 2 π-bond

Two sp-Hybrids orbitals





♦ C-H: σ-Bond (sp-s)

♦ C-C: σ-Bond (sp- sp)

♦ C=C: π-Bond (P_z- P_z)
(P_y- P_y)

$\underbrace{CH{\equiv}CH}$

♣ Effective electron gp= 2

♣ sp-hybridization

♣ Linear geometry and shape

♣ Bond angle = 180^0

Quiz

♦ Predict the Hybridization, geometry and shape :

A) NO_3^-

B) ClO_3^-

C) ClF_3

D) IF_5