KIM1170
GENERAL CHEMISTRY
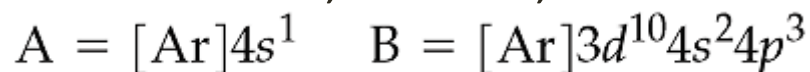
*Practices on Chemical Bonds
*Liquids and Intermolecular Forces

Practices

1. Two elements, A and B, have the electron configurations shown.

$$A = [Ar]4s^1 \quad B = [Ar]3d^{10}4s^24p^3$$

(a) Which element is a metal?

(b) Which element has the greater ionization energy?

(c) Which element has the larger atomic radius?

(d) Which element has the greater electron affinity?

A = 4. Period 1A Group   B = 4. Period 5A Group

(a) Metal

(b) In the same period → Ionization Energy

4 → 1A × 3A < 2A < 4A < 6A < 5A < 7A

B is the greater Ionization Energy. (d) Also electron affinity

(c) In the same period → Atomic Radius (-) → decrease

A is the larger atomic radius / diameter.

2. ($_{13}$Al, $_8$O, $_{11}$Na, $_{16}$S, $_{12}$Mg, $_7$N) --order the elements by increasing electron affinities.

Period		
3	$_{13}$Al : $1s^2 2s^2 2p^6 (3s^2 3p^1$	3. Period 3A Group
2	$_8$O : $1s^2 (2s^2 2p^4$	2. Period 6A Group
3	$_{11}$Na : $1s^2 2s^2 2p^6 (3s^1$	3. Period 1A Group
3	$_{16}$S : $1s^2 2s^2 2p^6 (3s^2 3p^4$	3. Period 6A Group
3	$_{12}$Mg : $1s^2 2s^2 2p^6 (3s^2$	3. Period 2A Group
2	$_7$N : $1s^2 2s^2 2p^3$	2. Period 5A Group

Electron affinity

$\xrightarrow{}$ (+)
$\xleftarrow{}$ (−)
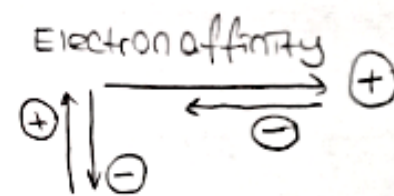
3. Period

1A	2A	3A	6A
Na	Mg	Al	S

2. Period

5A	6A
N	O

3

3. $(_{11}Na, _{12}Mg, _{12}Mg^{2+}, _7N^{3-}, _8O^{2-})$ --order the elements by increasing atomic radius.

$_{11}Na : 1s^2 2s^2 2p^6 3s^1$ 3. Period 1A Group

$_{12}Mg : 1s^2 2s^2 2p^6 3s^2$ 3. Period 2A Group $_{12}Mg$)) $\overline{12e^-}$

$_{12}Mg^{2+} : 1s^2 2s^2 2p^6 3s^2 \rightarrow$ $_{12}Mg$)) $\boxed{10e}$

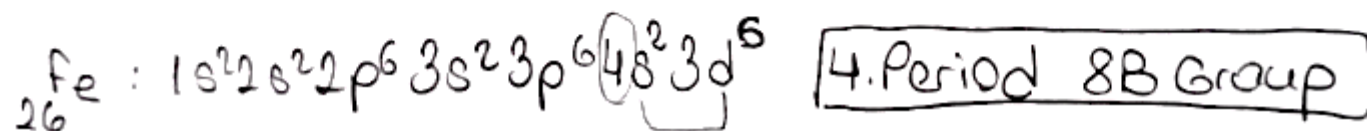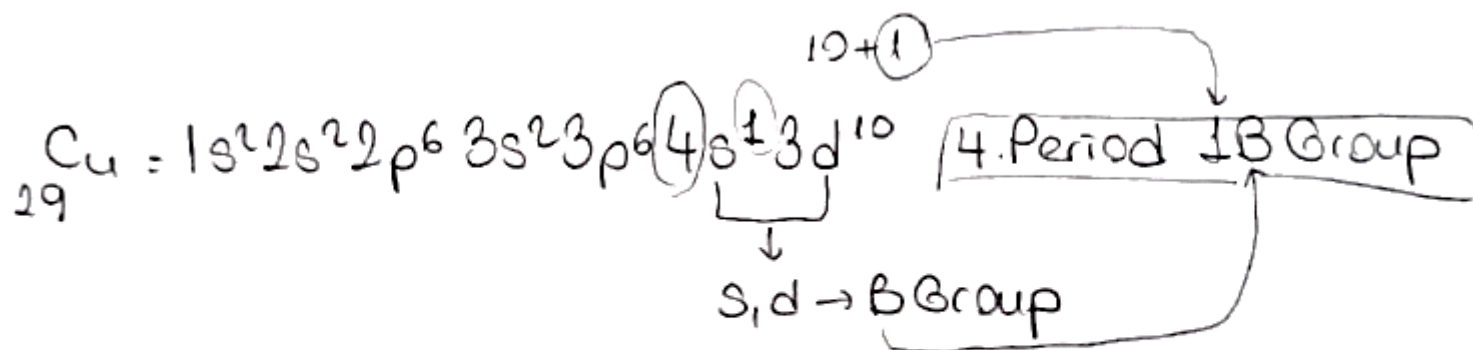The atomic volume is less suppressed in nitride ion because of the lower proton charge having 10 electron ions.

$_7N^{3-} : 1s^2 2s^2 2p^6 \rightarrow _7N^{3-}$)) $\boxed{10e^-}$

$_8O^{2-} : 1s^2 2s^2 2p^6 \rightarrow _8O^{2-}$)) $\boxed{10e}$

Atomic Radius = $(10e^-)$

$_{12}Mg^{2+} < _8O^{2-} < _7N^{3-} < _{12}Mg < _{11}Na$

2. Period   3. Period

4

4. Find the places of $_{29}Cu$ and $_{26}Fe$ elements in the periodic table.

$$_{29}Cu = 1s^2 2s^2 2p^6 3s^2 3p^6 (4)s^1 3d^{10} \quad \boxed{\text{4.Period 1B Group}}$$

$$10+\textcircled{1}$$

$$s, d \rightarrow B\,Group$$

$$_{26}Fe : 1s^2 2s^2 2p^6 3s^2 3p^6 (4)s^2 3d^6 \quad \boxed{\text{4.Period 8B Group}}$$

Practices:

1. Write Lewis symbols for the following elements.
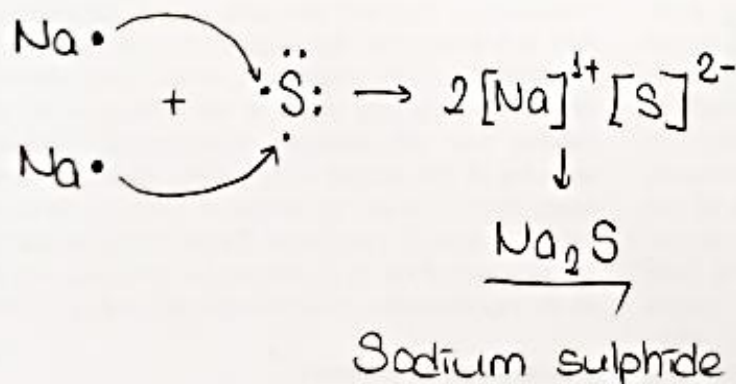
$_{12}Mg$, $_{32}Ge$, $_{19}K$, $_{10}Ne$, $_{50}Sn$, $_{35}Br^{-}$, $_{16}S^{-2}$

Also in a pratic way; the proton numbers with nobble gases are regulated in the periodic table as: $2 \rightarrow He \rightarrow 1.P$

$10 \rightarrow Ne \rightarrow 2.P$

$11 \,(12) \qquad (16) \, 17 \, 18 \rightarrow Ar \rightarrow 3.P$

$(19) \qquad 32 \, 33 \, 34 \, 35 \, (36) \rightarrow Kr \rightarrow 4.P$

$(50) \, 51 \, 52 \, 53 \, 54 \rightarrow Xe \rightarrow 5P$

$(P = Period)$

• Mg • $Mg \rightarrow 3.Period \, (2A \, Group$
            $_{12}$

• Ge • $Ge \rightarrow 4.Period \, (4A \, Group$
            $_{32}$

K • $K \rightarrow 4.Period \, (1A \, Group$
         $_{19}$

: Ne: $Ne \rightarrow 2.Period \, (8A \, Group$
         $_{10}$

• Sn • $_{50}Sn \rightarrow 5.Period \, (4A \, Group$

$[:Br:]^{1-}$ $_{35}Br \rightarrow 4.Period \, (7A \, Group + (1)e^{-}$

$(4A) \quad 5A \, 6A \, 7A \quad 8A$

$_{16}S^{2-} \rightarrow 3.Period \, (6A \, Group + (2)e^{-} \; [:S:]^{2-}$

- Write Lewis structures for the following compounds.

($_{13}$Al, $_8$O, $_{11}$Na, $_{16}$S, $_{12}$Mg, $_7$N) Al_2O_3, Na_2S, Mg_3N_2

Differences among them

1.P → 2 ⟩ 8
2.P → 10 ⟩ 8
3.P → 18 ⟩ 18
4.P → 36 ⟩ 18
5.P → 54 ⟩ 32
6.P → 86
↓
8A Group

$_{13}$Al → 3.Period 3A Group

$_8$O → 2.Period 6A Group

$_{11}$Na → 3.Period 1A Group

$_{16}$S → 3.Period 6A Group

$_{12}$Mg → 3.Period 2A Group

$_7$N → 2.Period 5A Group

$Al \cdot$ $\cdot \ddot{O} :$

$Al \cdot$ $+$ $\ddot{O} :$ $→ 2[Al]^{3+}3[O]^{2-}$

$\ddot{O} :$

↓

Al_2O_3

Aluminium Oxide

$Na \cdot$
$+$ $\cdot \ddot{S} : → 2[Na]^{1+}[S]^{2-}$
$Na \cdot$
↓
Na_2S

Sodium sulphide

$\cdot Mg \cdot$ $\cdot \ddot{N} \cdot$

$\cdot Mg \cdot$ $+$ $→ 3[Mg]^{2+}2[N]^{3-}$

$\cdot Mg \cdot$ $\cdot \ddot{N} \cdot$

↓

Mg_3N_2
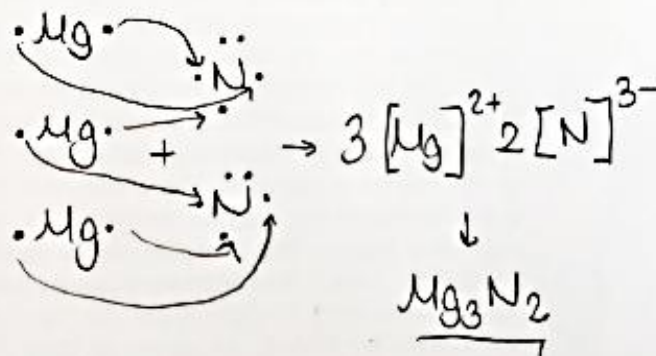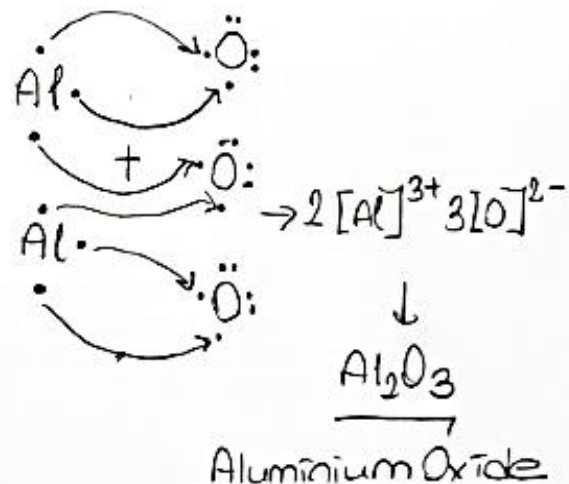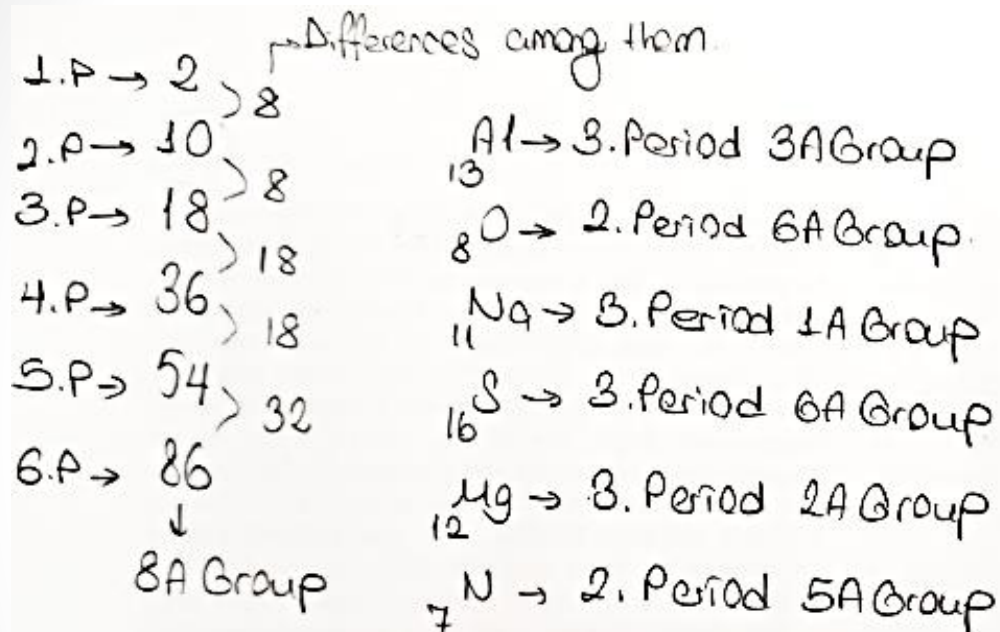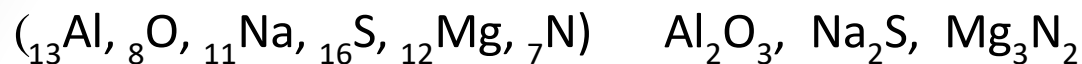
Lewis Structures and Formal Charge

- Therefore, the following steps are taken into account:

1. Determine the total number of valence electrons that must appear in the structure. Examples: C atom has 4 valence electrons

In the polyatomic ion $PO_4{}^{3-}$ there are 5 valence electrons for the P atom and 6 for each O atom. To produce the charge of an additional 3 valence electrons must be brought into the structure. So the total number of valence electrons in the Lewis structure of is;

$$5 + 24 + 3 = 32$$

In the polyatomic ion $NH_4{}^{+}$ there are 5 valence electrons for the N atom and 1 for each H atom. To account for the charge of one of the electrons must be lost. So the total number of valence electrons in is;

$$5 + 4 - 1 = 8$$

2. Identify the central atoms(s) and terminal atoms. **The atom having the most lowest electronegativity is the central atom.**

3. Write a plausible skeletal structure. Join the atoms in the skeletal structure by single covalent bonds (single dashes, representing two electrons each).

4. For each bond in the skeletal structure, subtract two from the total number of valence electrons.

5. With the valence electrons remaining, first complete the octets of the terminal atoms. Then, to the extent possible, complete the octets of the central atom(s). If there are just enough valence electrons to complete octets for all the atoms, the structure at this point is a satisfactory Lewis structure.

6. If one or more central atoms are left with an incomplete octet after step 5, move lone-pair electrons from one or more terminal atoms to form multiple covalent bonds to central atoms. Do this to the extent necessary to give all atoms complete octets, thereby producing a plausible Lewis structure.

7. Determine the formal charge of each atom in the structure.

- **Write a plausible Lewis structure for cyanogen, C₂N₂**

1. The total number of valence electrons is 4 + 4 + 5 + 5 = 18.
2. According to periodic table, C (4A) atoms have a lower electronegativity than the N (5A) atoms, C atoms are central atoms, and N atoms are terminal atoms.
3. A plausible skeletal structure by joining atoms through single covalent bonds
   N¬C¬C¬N
4. Subtract two electrons for each bond in the skeletal structure. The three bonds in this structure account for 6 of the 18 valence electrons. This leaves 12 valence electrons to be assigned.
5. Complete octets for the terminal N atoms, and to the extent possible, the central C atoms. The remaining 12 valence electrons are sufficient only to complete the octets of the N atoms.

$$:\ddot{N}\!-\!C\!-\!C\!-\!\ddot{N}:$$
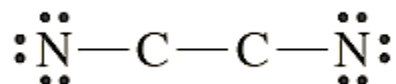
- Move lone pairs of electrons from the terminal N atoms to form multiple bonds to the central C atoms. Each C atom has only four electrons in its valence shell and needs four more to complete an octet. Thus, each C atom requires two additional pairs of electrons, which it acquires if we move two lone pairs from each N atom into its bond with a C atom, as shown below.

$$:\ddot{N}\!\!-\!\!C\!-\!C\!\!-\!\!\ddot{N}:$$

$$\downarrow$$

$$:N\equiv C\!-\!C\equiv N:$$
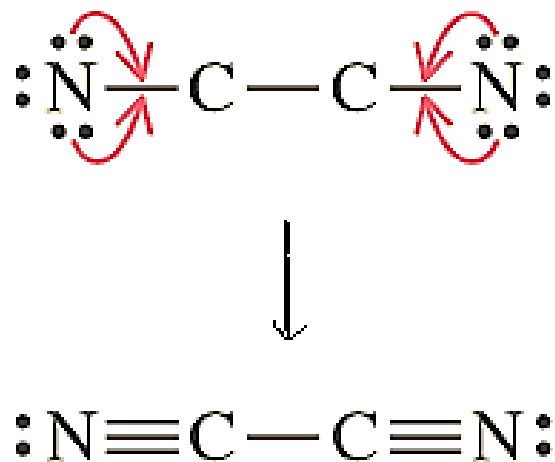
$$FC = \text{number valence } e^- \text{ in free atom} - \text{number lone-pair } e^- - \frac{1}{2}\text{number bond-pair } e^- \text{ (number of bonds made by the atom)}$$

(Formal Charge)

:N≡ =5 valence e⁻ - 2 lone-pair e⁻ - 1/2 x 6 bond-pair e⁻ → 0

≡C— =4 valence e⁻ - 0 lone-pair e⁻ - 1/2 x 8 bond-pair e⁻ → 0

:N≡C—C≡N:

- Write the Lewis structures for the covalent bonded compounds in below.

CH_3COO^-, SO_4^{2-}, H_2SO_4, CO_3^{2-}, N_2O, O_3, HCO_2^-

CH_3COO^-
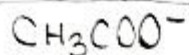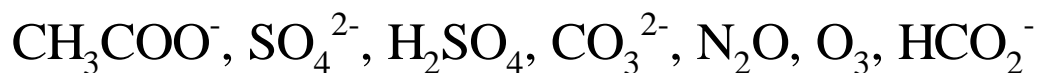
$_6C : 1s^2 2s^2 2p^2$ 2.Period ④A Group

$_1H : 1.$ Period ① A Group.

$_8O : 1s^2 2s^2 2p^4$ 2.Period ⑥A Group

① $2C + 3H + 2O + 1e^-$

$\rightarrow (2 \cdot 4) + (3 \cdot 1) + (2 \cdot 6) + 1$

$\rightarrow 8 + 3 + 12 + 1$

$\rightarrow 24e^-$ } The total valence e⁻ number that must appear in the lewis structure.

② Central atom is determined.
Electronagativity → ④A < 6A ✓

③ It is consisted of a plausible skeletal structure with the single covalent bonds firstly:

H—C—C—Ö: with H, :Ö: above C, H below C
Not comly with the octet rule

→ $\left[H-C-C=\overset{..}{O} \right]^-$ resonance form $\left[H-\overset{H}{\underset{H}{C}}-\overset{O}{C}-\overset{..}{O}: \right]^{-}$

$H = 1-0-1 = 0$
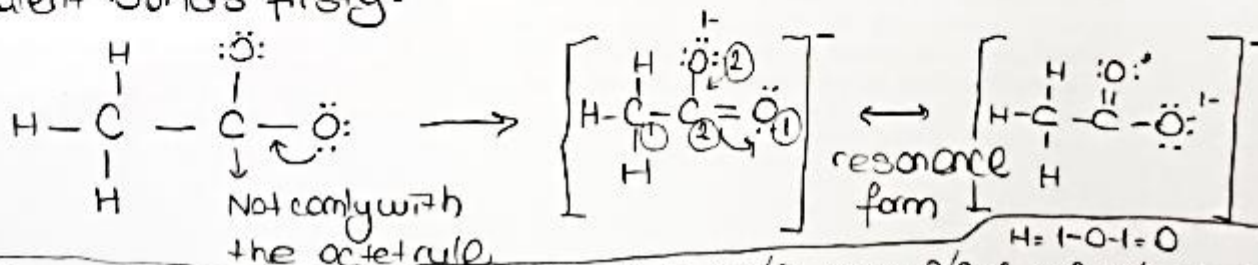
④ Determine the formal charge → $C_1: 4-0-4-0/C_2: 4-0-4=0/O_1: 6-4-2=0/O_2: 6-6 4=$

13

SO_4^{2-}
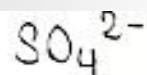
$_{16}S : 1s^2 2s^2 2p^6 3s^2 3p^4$ 3.Period 6A Group

$_8O : 1s^2 2s^2 2p^4 \rightarrow$ 2.Period | 6A Group

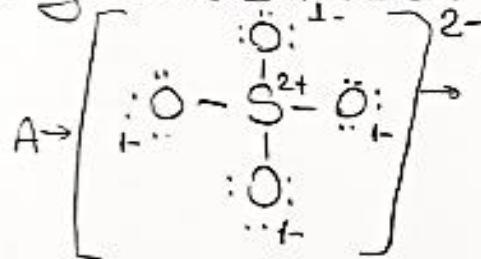② $S \rightarrow$ Central atom

③ Single covalent bonds formation

① Total VEN

$(S \times 1) + (O \times 4) + 2$

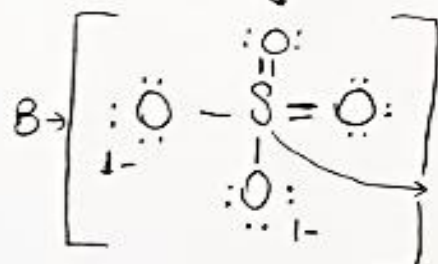$\rightarrow (6 \times 1) + (6 \times 4) + 2$

$\rightarrow 6 + 24 + 2 = 32 e^-$

A →

B →

The whole atoms are compatible with the octet rule. But the charge of the molecule ion is (2−).

So we can consist of π bonds, because

The formal charges for S/O atom is:

$O \rightarrow 6 - \quad 6 \quad - \quad 1 \quad = -1$

$\underset{VEN}{} \quad \underset{lone\ pair\ e^-}{} \quad \underset{bond\ number}{}$
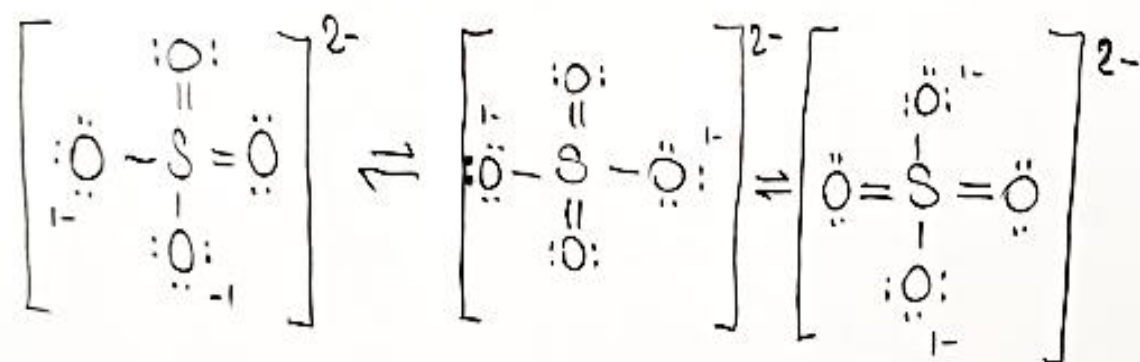
$S \rightarrow 6 - \quad 0 - 4 = +2$

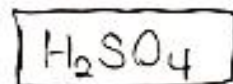$S_{formal\ charge} = 6 - 6 = 0$

→ Because of the small number of negative charges, the lewis structure B is more stable compared with A

14

And we can write resonance forms of B structure:

$$\left[\begin{array}{c} :\overset{..}{O}: \\ \| \\ :\overset{..}{\underset{..}{O}}{}^{1-} \sim S = \overset{..}{\underset{..}{O}} \\ | \\ :\overset{..}{\underset{..}{O}}:{}_{-1} \end{array}\right]^{2-} \rightleftharpoons \left[\begin{array}{c} :\overset{..}{O}: \\ \| \\ {}^{1-}:\overset{..}{\underset{..}{O}} - S - \overset{..}{\underset{..}{O}}:{}^{1-} \\ \| \\ :\overset{..}{O}: \end{array}\right]^{2-} \rightleftharpoons \left[\begin{array}{c} :\overset{..}{\underset{|}{O}}:{}^{1-} \\ \overset{..}{O} = S = \overset{..}{O} \\ | \\ :\overset{..}{\underset{..}{O}}:{}_{1-} \end{array}\right]^{2-}$$

*we prefer these resonance forms, although the central S atom exceed the octet rule. But the resonance forms are determined considering the net charge of the molecule.

H_2SO_4

$_1H \to 1.$ Period 1A Group
$_{16}S \to 3.$ Period 6A Group
$_8O \to 2.$ Period 6A Group

2
10
18
36
54

Central atom.

$VEN \Rightarrow (2 \times H) + (S) + (4 \times O)$
$\rightarrow (2 \times 1) + 6 + (4 \times 6)$
$\rightarrow 32e^-$

H
|
:O:
|
:O̤ — S — O̤ — H
|
:O̤:
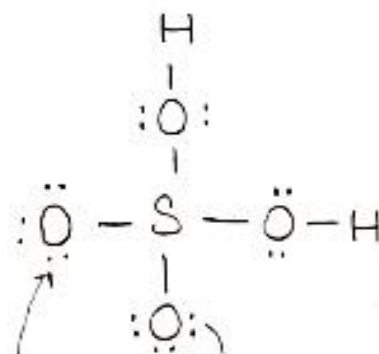
$12 e^- \to$ electron numbers
participated to chemical bond

$32 - 12 = 20 e^- \to$ lone pair electrons.
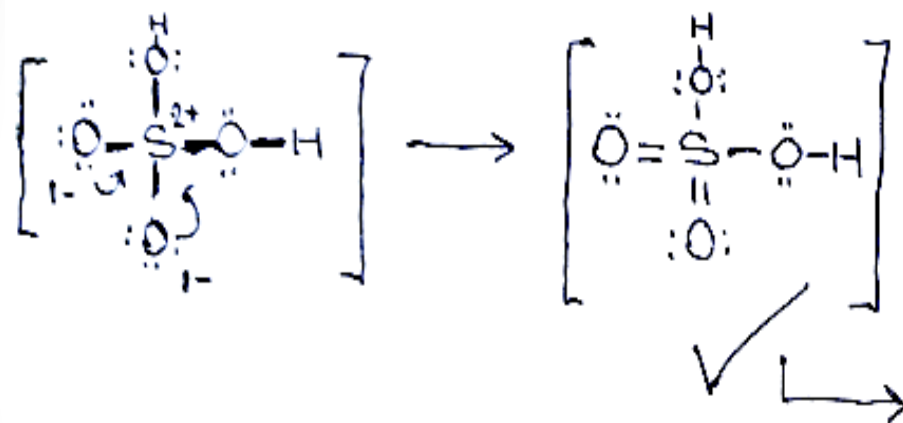
$\rightarrow$ Finally, if we look the formal charges;

:O̤ — $\rightarrow$ $6 - 6 - 1 = -1$
             VEN free e⁻ bond number
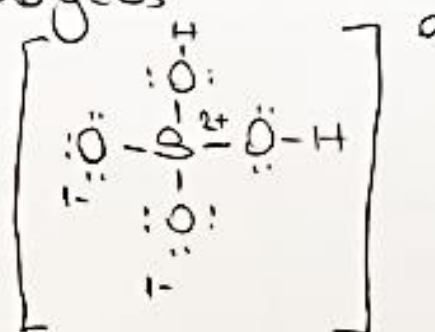
$-S- \rightarrow 6 - 0 - 4 = +2$

$-O̤- \rightarrow 6 - 4 - 2 = 0$

$-H \rightarrow 1 - 0 - 1 = 0$

$$\left[\begin{array}{c} H \\ | \\ :O̤: \\ | \\ :O̤ - S^{2+} - O̤ - H \\ | \\ :O̤: \end{array} \right]^0$$

16

$$\ddot{O} = \left(\begin{array}{c}\text{formal}\\\text{charge}\end{array}\right) = 6 - 4 - 2 = 0$$

$$= S - \left(\begin{array}{c}\text{formal}\\\text{charge}\end{array}\right) = 6 - 0 - 6 = 0$$

$$-H \left(\begin{array}{c}\text{formal}\\\text{charge}\end{array}\right) = 1 - 0 - 1 = 0$$

$$-\ddot{O} - \left(\begin{array}{c}\text{formal}\\\text{charge}\end{array}\right) = 6 - 4 - 2 = 0$$

CO_3^{2-}
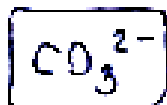
$_6C : 1s^2 2s^2 2p^2$ ② Period 4A Group

$_8O : 1s^2 2s^2 2p^4$ ② Period 6A Group

① $1C + 3 \cdot O + 2e^-$

$\rightarrow (1 \times 4) + (3 \times 6) + 2e^-$

$\rightarrow 5 + 18 + 2 = 25e^-$ Total VEN
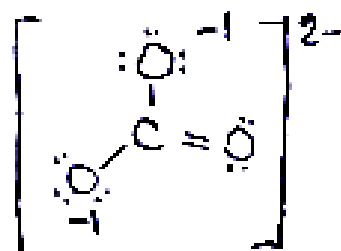
② Central Atom: C

Electronegativity: 4A < 6A

③
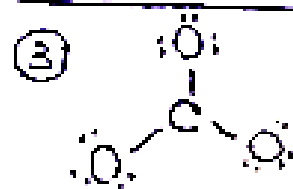
$25e^- \rightarrow$ Total VEN

$\rightarrow$ Plausible skeletal structure

$6e^- \rightarrow$ Bonding e^-

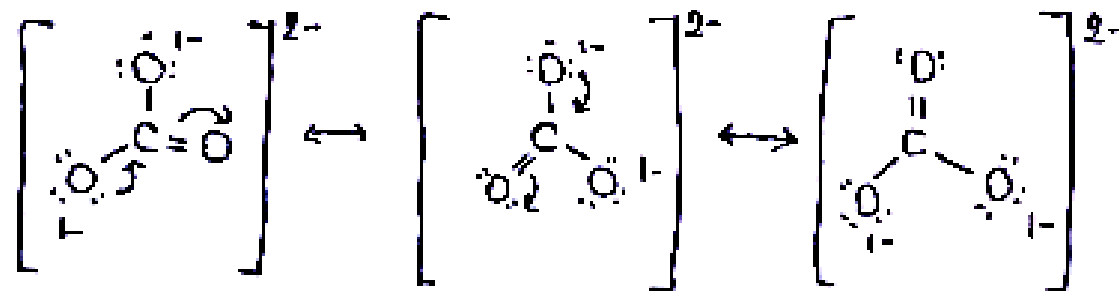$25-6 = 18e^- \rightarrow$ free e^-

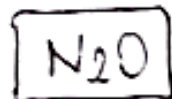④ Compare to the octet rule.

⑤ Calculate the formula charge.

$:\overset{..}{O}^- \rightarrow 6 - 6 - 1 = -1$

$\searrow C = \rightarrow 4 - 0 - 4 = 0$

$= \overset{..}{O} \rightarrow 6 - 4 - 2 = 0$

⑥ Resonance forms:

N_2O

$_7N: 1s^2 2s^2 2p^3$ / 2. Period ⑤A Group

$_8O: 1s^2 2s^2 2p^4$ / 2. Period ⑥A Group

Valence e⁻ number

① → $(2 \times N) + (1 \times O)$
→ $(2 \times 5) + (1 \times 6)$
→ $16e^- →$ Total VEN

② Electronegativity $N < O$
⑤A ⑥A
N → Central atom ✓

③ Single bonds are placed, firstly.

$:\ddot{N} - \ddot{N} - \ddot{\ddot{O}}:$ $16 - 4 = 12e^-$
VEN
• → excess $12e^-$ free e⁻
Bonding e⁻

$:\ddot{N} - \ddot{N} = \ddot{O}$

19
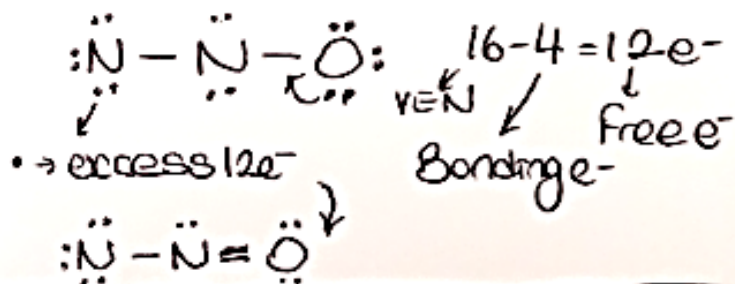
④ formal charges are calculated:

:N̈—N̈=Ö̈   :N̈—→   VEN — free e⁻ — bond number
(2-)   (0) ö̈   5 — 6 — 1   = (−2)

                 ⁻N̈= —→   5 — 2 — 3 = 0

The net charge   =Ö̈ —→   6 — 4 − 2 = 0
of the molecule
is (2−). So this is not comply with the natr form.

⑤ So try different structures considering the formal charges and
                                                    octet rule. **

(5.1)   N̈=N̈=Ö̈ → (6−4−2)=0   (5.2)   :N̈⁻²—N̈=Ö:¹⁺ → (6−2−3)=(+1)

(5−4−2)(−1)   (5−0−4)=(+1)   (5−6−1)(−2)   (5−0−4)=(+1)
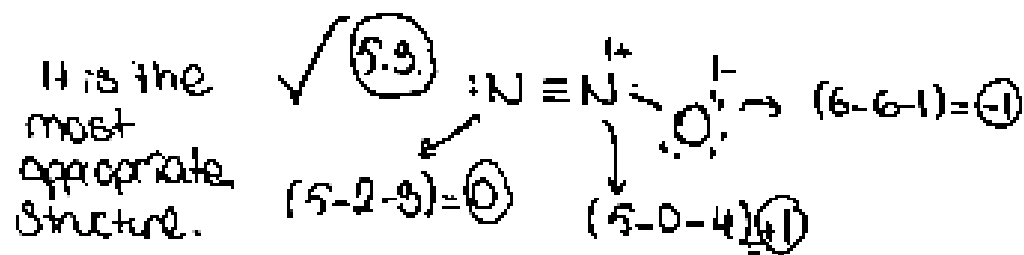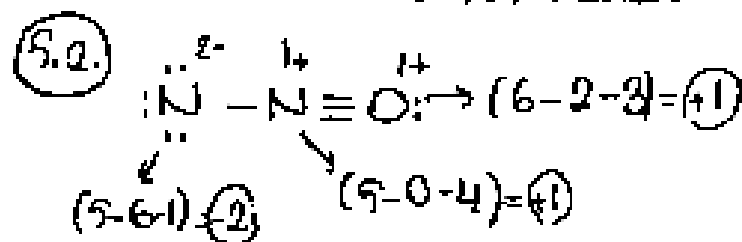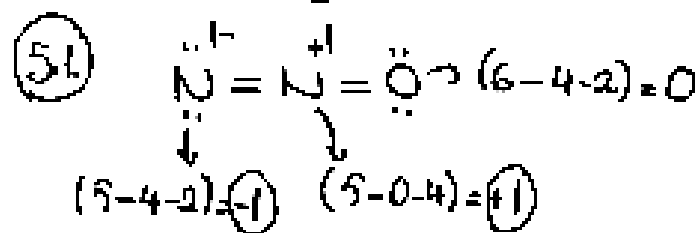
It is the √(5.3)   :N≡N̈⁻...O̅: → (6−6−1)=(−1)   Not: O atom is more
most   electronegative as we
appropriate   (5−2−3)=(0)   (5−0−4)(+1)   show the period and
structure.   group number: when
                                                        compared to "N" atom.

So 'O' atom takes negative charge in the lewis form
although the tightly triple bond in 5.3.
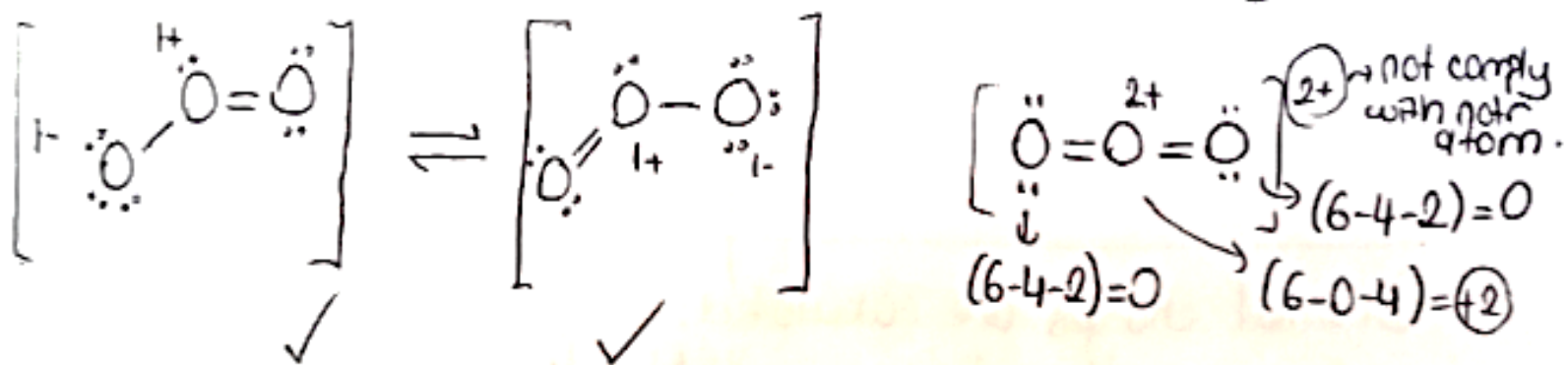
O_3 $_8O: 1s^2 2s^2 2p^4 / 2.Period$ ⑥A Group

① $3 \times 6 = 18 e^-$ VEN

② Central atom: O

③ $\overset{..}{O} - \overset{..}{\underset{..}{O}}:$ $18 - 4 = 14 e^-$ (free e)

$\overset{..}{\underset{..}{O}}$ ↳ excess 14

④ Determine the formal charges. $\longrightarrow$ $1 - \overset{..}{O} / \overset{\overset{..}{O}}{\underset{..}{O}} 1+ = \overset{..}{\underset{..}{O}} \rightarrow (6-4-2) = 0$

$(6-6-1) = \textcircled{-1}$ $(6-2-3) = \textcircled{+1}$

$\left[1- \overset{..}{\underset{..}{O}} / \overset{1+ \overset{..}{O} = \overset{..}{\underset{..}{O}}}{} \right] \rightleftharpoons \left[\overset{..}{O} / \overset{\overset{..}{O} - \overset{..}{O}:}{1+ \quad 1-} \right]$ ✓ ✓

$\left[\overset{..}{O} = \overset{2+}{O} = \overset{..}{\underset{..}{O}} \right]$ ⟶ ② 2+ → not comply with octet atom.

$(6-4-2) = 0$ $(6-4-2) = 0$ $(6-0-4) = \textcircled{+2}$

21

$\boxed{HCO_2^-}$ Formate ion

$_6C$: 2. Period $\textcircled{4}A$ Group

$_8O$: 2. Period $\textcircled{6}A$ Group

① $(1 \cdot H) + (1 \cdot C) + (2 \cdot O) + 1e^-$
→ $(1 \cdot 1) + (1 \times 4) + (2 \cdot 6) + 1$
→ $1 + 4 + 12 + 1$
→ $18e^- →$ total VEN.

② Electronegativity $4A < 6A$
$\quad\quad\quad\quad\quad\quad\quad\quad\quad \downarrow \quad \downarrow$
$\quad\quad\quad\quad\quad\quad\quad\quad C < O$

Central atom → C

③ Single bonds and complete octet/
$\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad$ dublet(H)
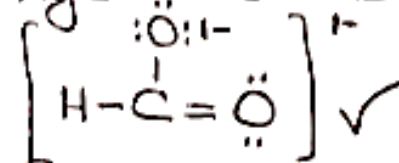
$$\left[\begin{array}{c} :\overset{..}{O}:^{1-} \\ | \\ H-C=\overset{..}{\underset{..}{O}} \end{array} \right]^{1-} \checkmark$$

④ Formal charges

$H- → (1-0-1)=0$
$-\overset{|}{C}= → (4-0-4)=0$

$:\overset{..}{O}:→ (6-6-1)=\boxed{-1}$
$=\overset{..}{\underset{..}{O}} → (6-4-2)=0$

⑤ Resonance form:

$$\left[\begin{array}{c} :\overset{..}{O}:^{1-} \\ | \\ H-C=\overset{..}{\underset{..}{O}} \end{array} \right]^{1-} \rightleftharpoons \left[\begin{array}{c} :O: \\ || \\ H-C-\overset{..}{\underset{..}{O}}:^{1-} \end{array} \right]^{1-}$$
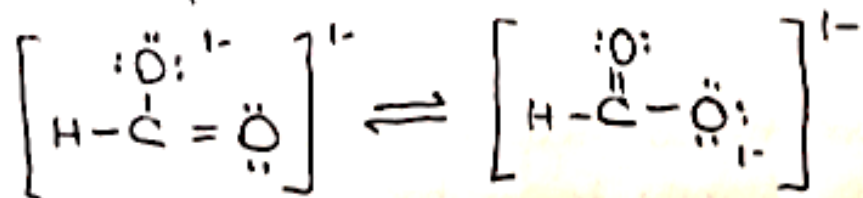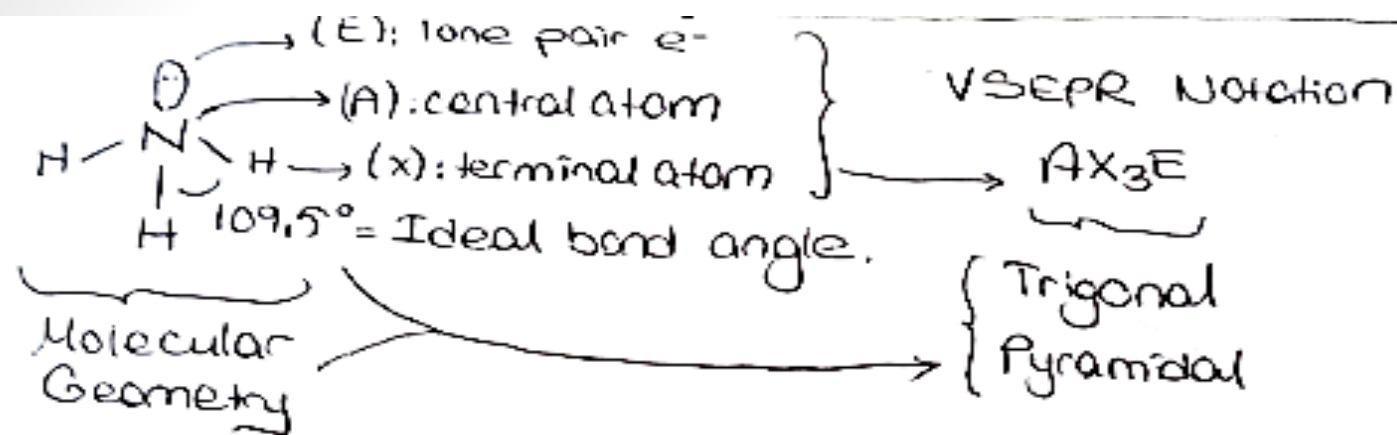
- Draw the shape of molecule geometries of the compounds in below, considering the VSEPR Theory. $_9F$, $_{53}I$, $_{54}Xe$, $_{17}Cl$)

$$NH_3, \; H_2O, \; PCl_5, \; SF_4, \; SF_6, \; IF_4^-, \; XeF_4$$

$\boxed{NH_3 = Ammonia}$

$_7N : 1s^2 2s^2 2p^3$ — electrons in half full orbitals which are used in covalent bonds.

lone pairs ← $\uparrow\downarrow$ ($2s$) $\boxed{\uparrow \;\; \uparrow \;\; \uparrow \atop P_x \;\; P_y \;\; P_z}$ $2p$

electrons (free)

sp^3 hybridization → hybrid orbital type → sp^3

1 unit s orbital ($2s$)

3 units co-energy orbitals. ($2p$) $P_x \; P_y \; P_z$

$\longrightarrow$ (E): lone pair e^-

$\longrightarrow$ (A): central atom

(diagram: N with lone pair, bonded to H, H, H)

$H-N$ $\backslash H \longrightarrow$ (x): terminal atom

H $109.5° =$ Ideal bond angle.
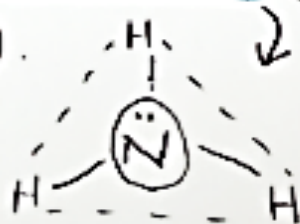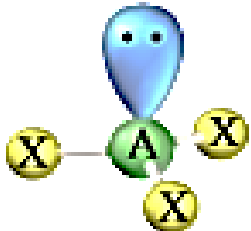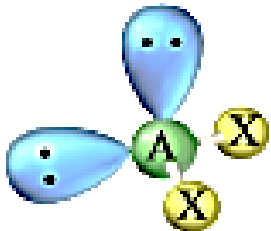
VSEPR Notation
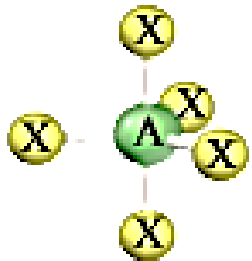
AX_3E

Molecular Geometry

Trigonal Pyramidal

$\ast$ Electron-Group Geometry $\longrightarrow$ Tetrahedral
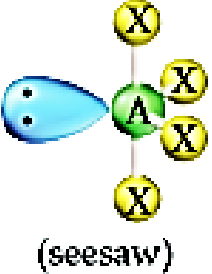
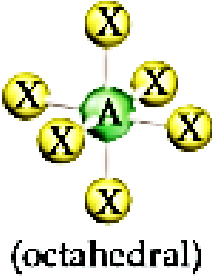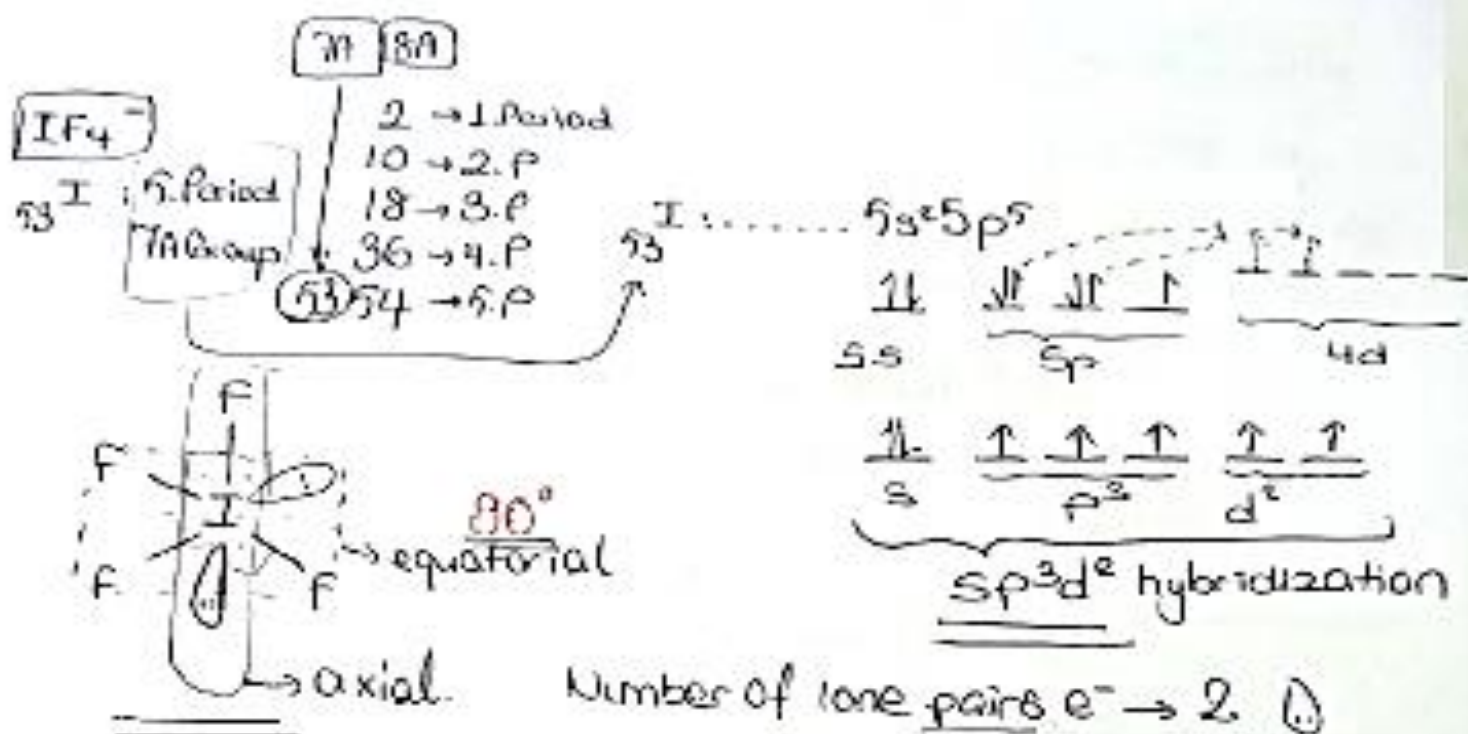When thought in three dimensions, $\longrightarrow$ so it can be called as tetrahedral for

4

e^- group geometry. But for molecular geometry it is appropriate to be called as trigonal pyramidal (In three dimension).
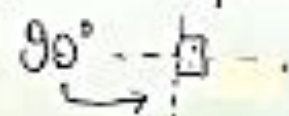
$\longrightarrow$ 1.

$\longrightarrow$ 2.

H 3. $\longleftarrow$ H

H $\longrightarrow$ 4. plane

(diagram: N with H, H, H in trigonal pyramidal arrangement)
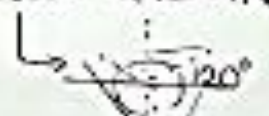
Hybride Type	Electron group geometry	Lone pairs electrons	VSEPR Notation	Molecular Geometry	Ideal bond angles	Molecule
sp3	tetrahedral	1	AX_3E	(trigonal pyramidal)	109.5°	NH_3
sp3	tetrahedral	2	AX_2E_2	(bent)	109.5°	OH_2
sp3d	trigonal bipyramidal	0	AX_5	(trigonal bipyramidal)	90°, 120°	PCl_5

Hybride Type	Electron-Group Geometry	Number of Lone Pairs	VSEPR Notation	Molecular Geometry	Ideal Bond Angles	Example
sp3d	trigonal bipyramidal	1	AX_4E^b	(seesaw)	90°, 120°	SF_4
sp3d2	octahedral	0	AX_6	(octahedral)	90°	SF_6
sp3d2	octahedral	2	AX_4E_2	(square planar)	90°	XeF_4

$[IF_4^-]$

$_{53}I$; 5. Period
7A Group

$\frac{7A}{18A}$

$2 \rightarrow 1.Period$
$10 \rightarrow 2.P$
$18 \rightarrow 3.P$
$36 \rightarrow 4.P$
$\textcircled{53}\ 54 \rightarrow 5.P$

$_{53}I \cdots\cdots\cdots 5s^2 5p^5$

5s   5p   4d

s p^3 d^2

sp^3d^2 hybridization

$80°$ → equatorial

→ axial

Number of lone pairs $e^- \rightarrow 2$

* The lone pairs electrons can be placed both equatorial and axial in this molecule geometry. Because the all angles are 90°.

* If the equatorial angles are 120° and the axial angles are 90° , , in a structure, so it can be prefered to be placed lonepairs electrons to equatorial firstly, because of the expanse area, so the prevention of steric hindrance.

120°