Chemical Bonding -1
Exercise - 1
PART - I

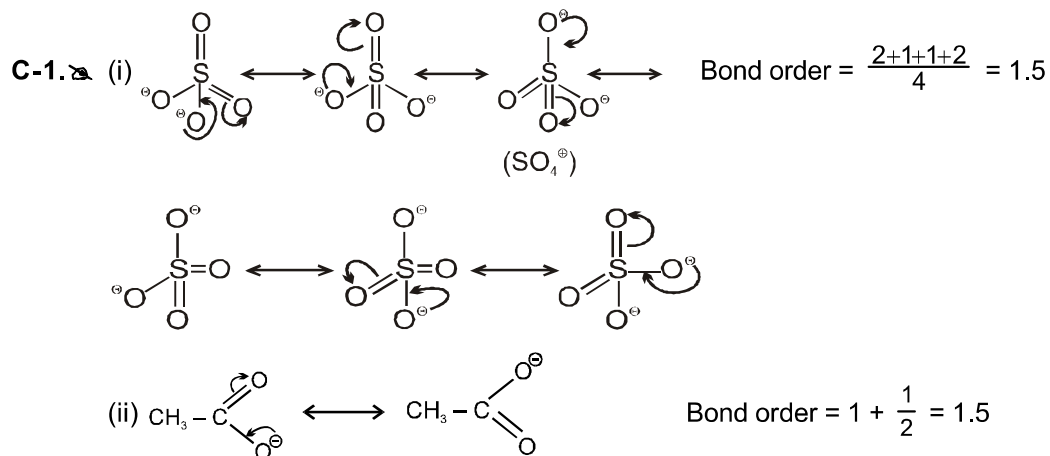**A-1.** (diagrams of Lewis structures: $F-Be-F$; $SiCl_4$; $H-C(=O)-O-H$; $H-O-S-O-H$ with two O; $O-O$; OF_2; $O=N-Cl$; $SOBr_2$; SO_2Cl_2)

A-2. (i) 4   (ii) 4   (iii) 3

A-3. (a) $[:\ddot{S}-C\equiv\ddot{N}]^-$ (b) Covalent compound (Cl_2O structure)
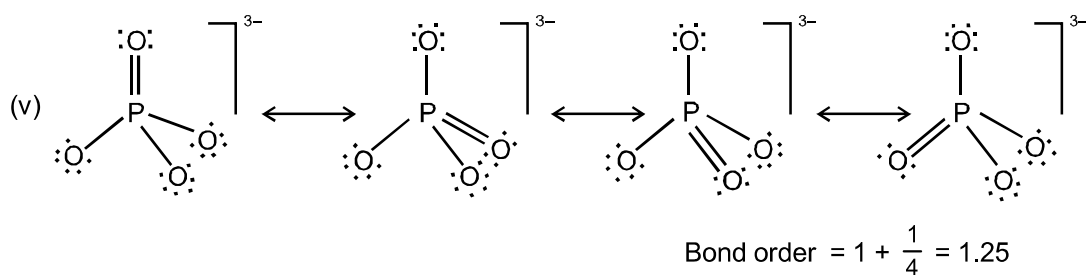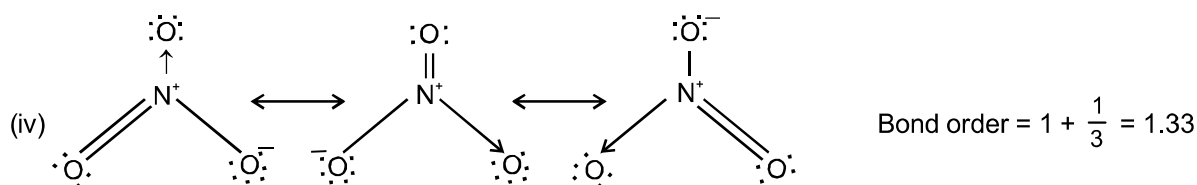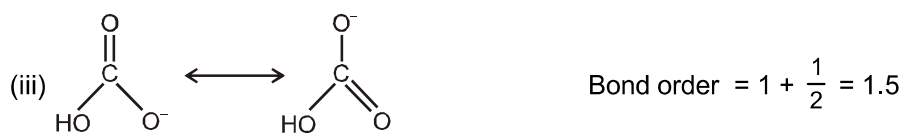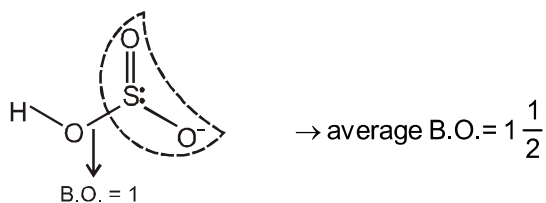
(c) (NO_2 structure)   (d) (NCl_3 structure)
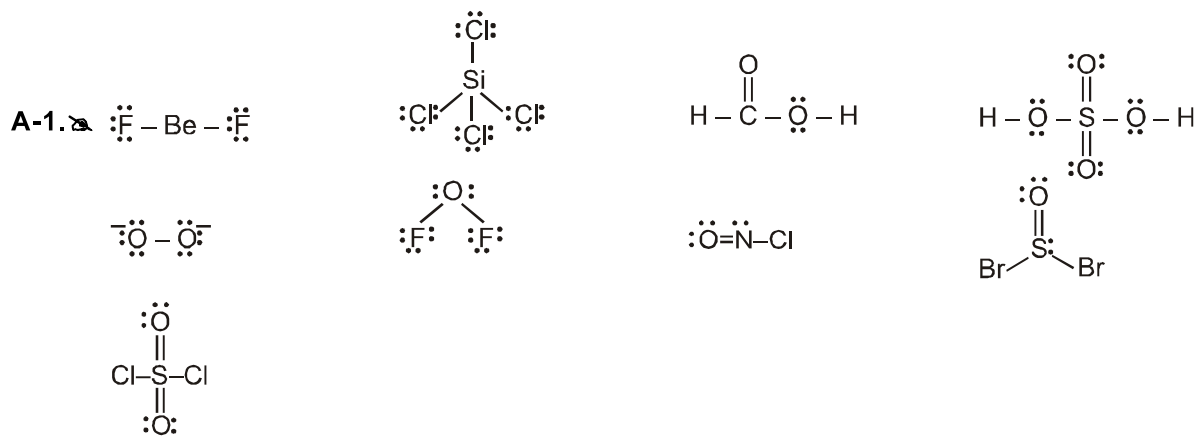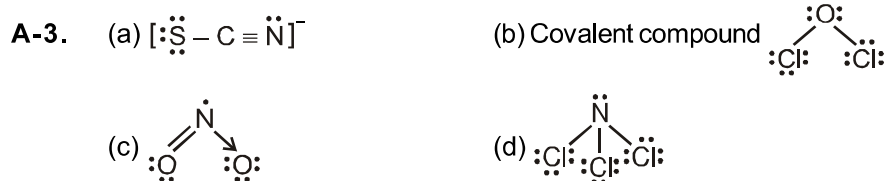
A-4. 5

B-1. **5**

B-2. BCl_3 electron deficient molecule
XeF_2 Super octet molecule
NO odd electron
IF_7 Super octet molecule
NO_2 odd electron
ClF_3 Super octet molecule
ClO_2 odd electron

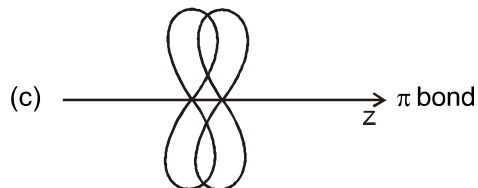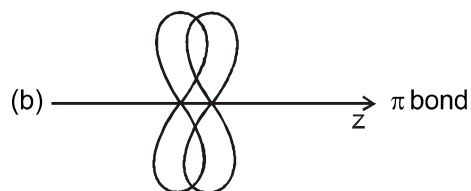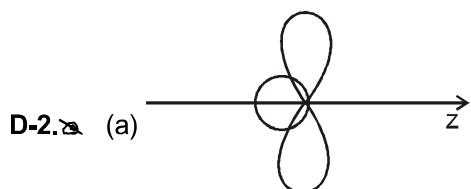**C-1.** (i) (resonance structures of SO_4^{2-})   Bond order = $\frac{2+1+1+2}{4}$ = 1.5

($SO_4^{\ominus}$)

(resonance structures)

(ii) $CH_3-C(=O)-O^{\ominus} \longleftrightarrow CH_3-C(-O^{\ominus})=O$ Bond order = $1 + \frac{1}{2}$ = 1.5

(iii) Bond order $= 1 + \dfrac{1}{2} = 1.5$

(iv) Bond order $= 1 + \dfrac{1}{3} = 1.33$

(v) Bond order $= 1 + \dfrac{1}{4} = 1.25$

C-2. (a) 2   (b) 1

C-3. Sodium formate exists as $HCOO^- Na^+$

C-4. $\rightarrow$ average B.O. $= 1\dfrac{1}{3}$

 $\rightarrow$ average B.O. $= 1\dfrac{1}{2}$

Hence one bond in HSO_3^- is longer than S–O bond in SO_3^{2-}. But other two S–O bond in HSO_3^- are shorter bonds.

D-1. Number of sigma bonds is 7 & number of pi bonds is 3.

D-2.

(a)

(b) π bond

(c) π bond

(d) σ bond

(e) z σ bond

(f) x + − + z zero overlap y

(g) y + + − z zero overlap

(h) − + + − + σ

(i) − + + − + z zero overlap
xy plane

(j) + + − x − + z zero overlap
xy plane

(k) + + + − Zero overlap

(l) x + + − + z zero overlap y

(m) x y z δ bond

(n) x z y δ bond d_{xy} d_{xy}

(o) x z = zero overlap y
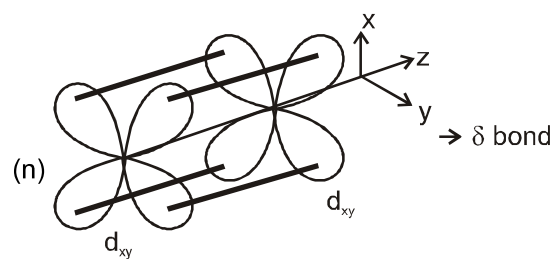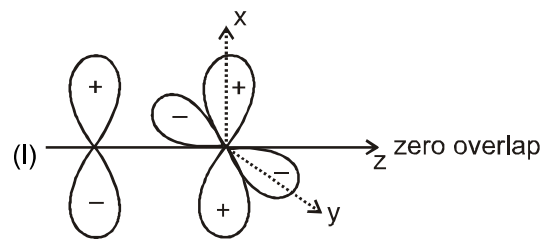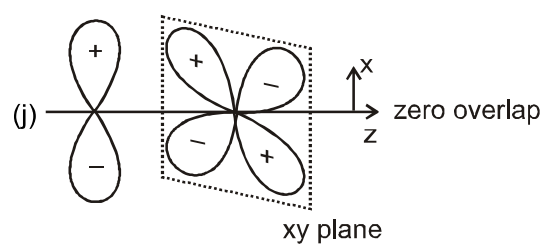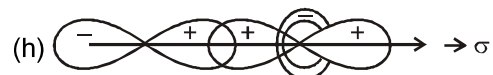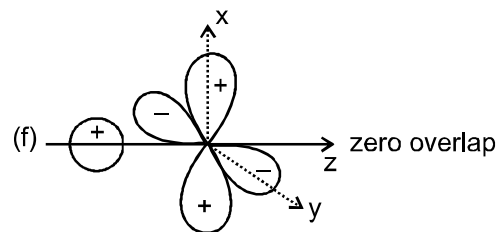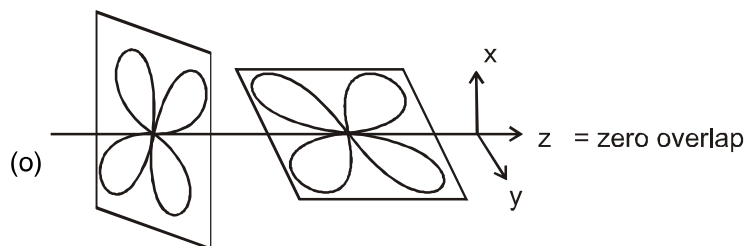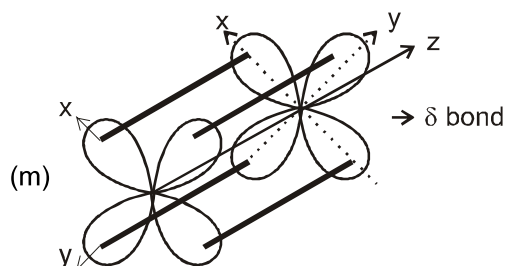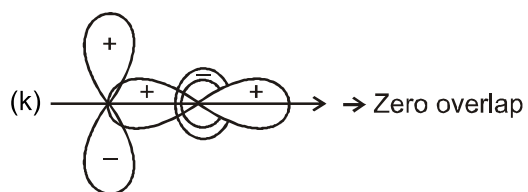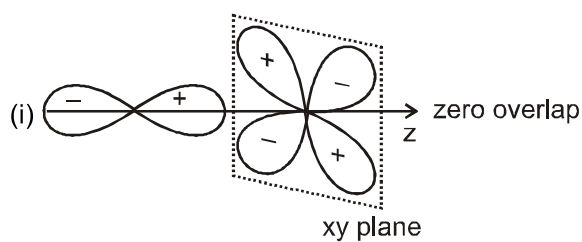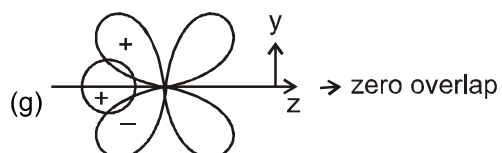
PART - II

A-1.	(C)	**A-2.**	(D)	**B-1.**	(D)	**B-2.**	(D)	**B-3.**	(A)	**B-4.**	(D)	**C-1.**	(C)
C-2.	(D)	**C-3.**	(D)	**C-4.**	(D)	**C-5.**	(C)	**C-6.**	(D)	**D-1.**	(C)	**D-2.**	(C)
D-3.	(B)	**D-4.**	(A)	**D-5.**	(A)								

PART - III

1. (A) → q, (B) → r, (C) → p, (D) → s

Exercise - 2
PART - I

1. (B) 2. (D) 3. (B) 4. (D) 5. (C) 6. (D) 7. (A)
8. (A) 9. (D) 10. (C)

PART - II

1. 2 2. 5 3. 10 4. 3 5. 12 6. 5
7. 5

PART - III

1. (ACD) 2. (AC) 3. (BC) 4. (AB) 5. (ABC)

PART - IV

1. (D) 2. (D) 3. (C) 4. (B) 5. (A)

Exercise 3
PART - I

1.* (ACD) 2. 8 3.* (BC) 4. 6

PART - II

1. (3) 2. (1)

Chemical Bonding-2
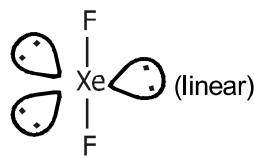## Exercise - 1
Part - I

A-1. According to VSEPR theory
XeF_2
Number of electron pairs = 5,
Number of bond pairs = 2,
So, Number of lone pairs = 3
Thus XeF_2 is linear with 3 lone pairs occupying
3 equatorial positions of trigonal bipyramidal so as to minimize the repulsions.

 (linear)

XeF_4
Number of electron pairs = 6,
Number of bond pairs = 4,
So, Number of lone pairs = 2
Thus XeF_4 is linear with 2 lone pairs occupying
2 axial positions of octahedral pyramidal so as to minimize the repulsions.

 (square planar)
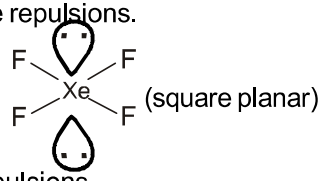
XeO_2F_2
Number of electron pairs (including super electron pairs) = 5,
Number of bond pairs = 4,
So, Number of lone pairs = 1
Thus XeO_2F_2 is see-saw with 1 lone pairs occupying
one equatorial position and two double bonds occupying other
two equatorial positions of trigonal bipyramidal so as to minimize the repulsions.

 (see-saw)
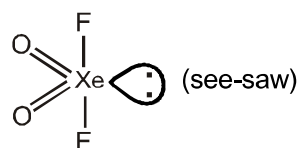
A-2. There are 5 electron pairs and all are bonds pairs in PCl_5. So to have the minimum repulsions between bond pairs it acquires trigonal bipyramidal shape. In BrF_5, there are 6 electrons pairs out of which one lone pair and rest all are bond pairs. So to have the minimum repulsions between bond pairs and lone pairs it acquires square pyramidal shape.
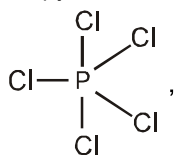
PCl_5 (trigonal bipyramidal), BrF_5 (square pyramidal)
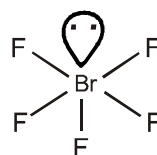
A-3. According to VSEPR theory two lone pairs out of six electron pairs are trans to each other to have minimum repulsion. The shape of XeF_4 is square planar and geometry is octahedral with sp^3d^2 hybridisation. The molecule looks like :

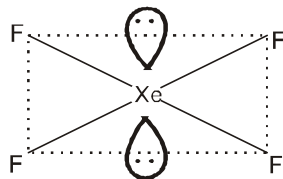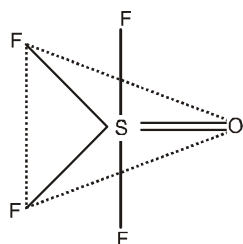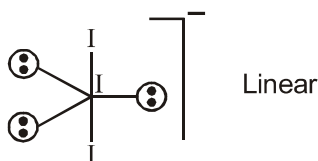In OSF_4, there are five electron pairs and all are bond pairs. So geometry is trigonal bipyramidal. As double bond creates more repulsion than singles bond, the double bond acquires one of equatorial position of trigonal bipyramidal to have minimum repulsions.
The structure looks like:

A-4. In NO_2^+ the N has sp hybridisation ; so it is linear $O = \overset{+}{N} = O$

In I_3^- there are 5 electron pairs around central iodine atom (3 lone pairs and 2 bond pairs). The hybridisation of iodine is thus sp^3d. To have minimum repulsions between lp-lp and lp-bp it acquires linear shape as shown below.

A-5. In PCl_5 there are 5 electron pairs around central phosphorus atom and all are bond pairs. The hybridisation of phosphorus is thus sp^3d. To have minimum repulsions between bp-bp it acquires trigonal bipyramidal shape as shown below.
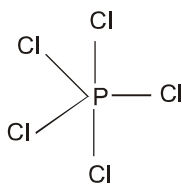
In IF_5 there are 6 electron pairs around central iodine atom. The hybridisation of iodine is thus sp^3d^2. 6 electron pairs contain 5 bond pairs and one lone pair so it will be square pyramidal to have minimum repulsions between lp-bp and bp-bp.

A-6.

nearly 'T' shaped.

B-1.

1. $F - \overset{\cdot\cdot}{\underset{\cdot\cdot}{Xe}} - F \longrightarrow sp^3d.$

2.

$\longrightarrow sp^3d^2$

3.

$\longrightarrow sp^3$

4.

$\longrightarrow sp^3d$

5.

$\longrightarrow sp^3d^2$

6. $IF_3 : -$

$\longrightarrow sp^3d$

7. $IF_5 : -$

$\longrightarrow sp^3d^2$

8. $IF_7 : -$

$\longrightarrow sp^3d^3$

9. $CH_4 : -$

$\longrightarrow sp^3$

10. $CCl_4 : -$

$\longrightarrow sp^3$

11. $SiCl_4 : -$

$\longrightarrow sp^3$

12. $SiH_4 : -$

$\longrightarrow sp^3$

13. $H_2O : -$

$\longrightarrow sp^3$

14. $\overset{\cdot\cdot}{N}H_3 \quad : -$

$\longrightarrow sp^3$

15.

16. BrF_5

$\longrightarrow sp^3d^2$

17. $O = N \underset{O^-}{\overset{O}{\diagup}}$ $\longrightarrow sp^2$ **18.** $O = C \underset{O^-}{\overset{O^-}{\diagup}}$ $\longrightarrow sp^2$

19. $H - \overset{+}{\underset{H}{N}} \overset{H}{\diagup} H$ $\longrightarrow sp^3$ **20.** $:Cl \overset{O^-}{\underset{O}{\equiv}} O$ $\longrightarrow sp^3$
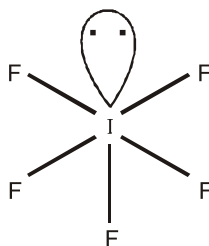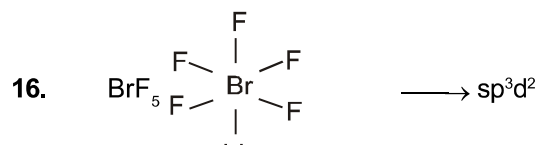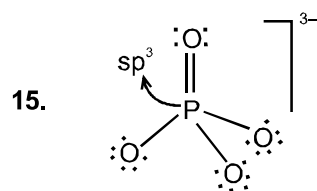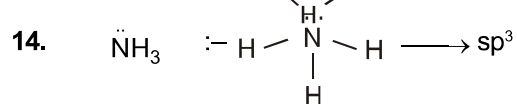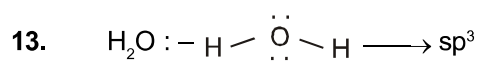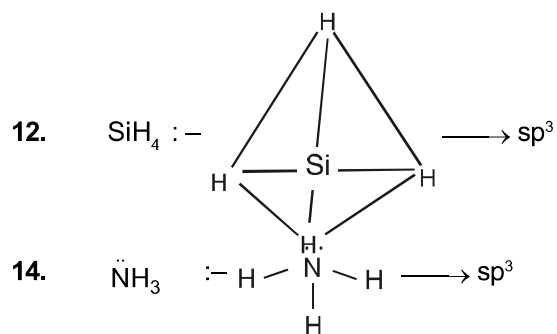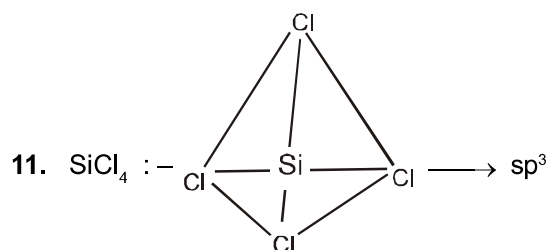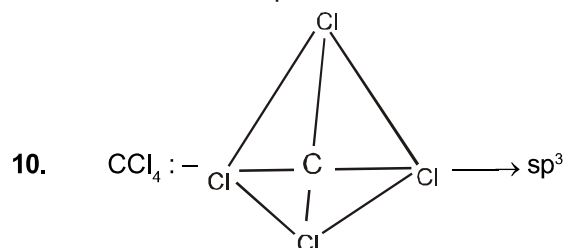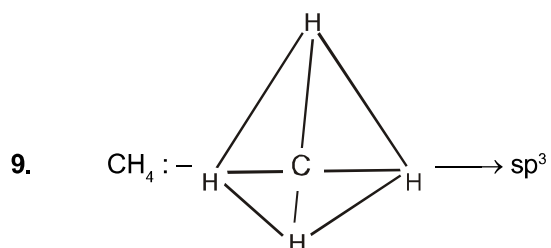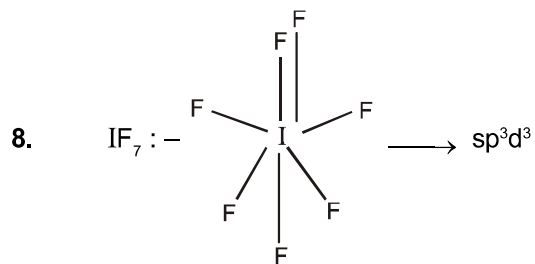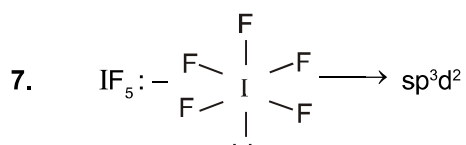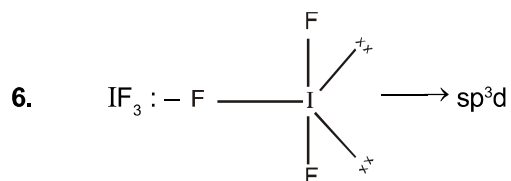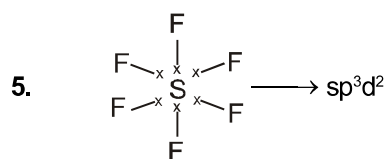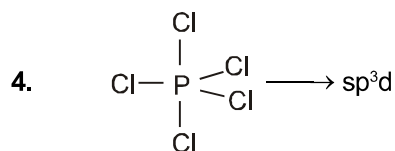
B-2. True, As the s character in hybrid orbital decrease, size of hybrid orbital increases.

B-3. (a) SF_4

4 bond pair & 1 lone pair, Hybridization = sp^3d Shape : see saw

(b) $XeOF_4$

5 bond pair & 1 lone pair, Hybridization = sp^3d^2 Shape : Square pyramidal

C-1. (a)

97.1°   **98.2°**   **99°**
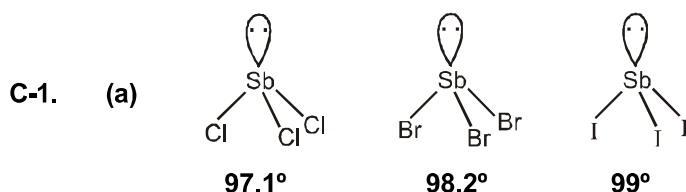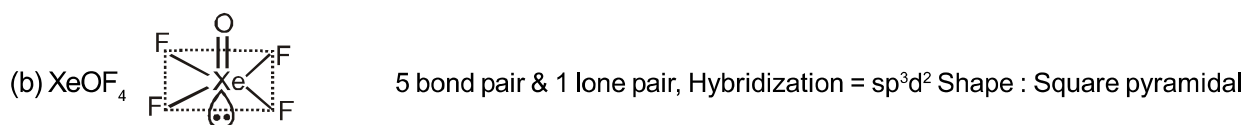
Cl, the most electronegative of the halogens in this series, pulls shared electrons the most strongly away from Sb, reducing electron density near Sb. The consequence is that the lone pair exerts the strongest influence on shape in $SbCl_3$.

(b)

102°   **100°**   **99°**
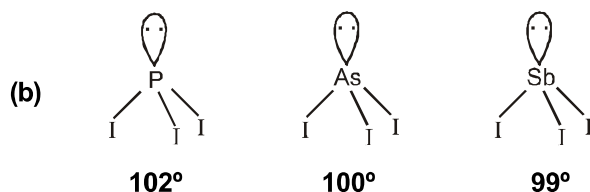
Phosphorus is the most electronegative of the central atoms. Consequently, it exerts the strongest pull on shared electrons, concentrating these electrons near P and increasing bonding pair-bonding pair repulsions– hence, the largest angle in PI_3. Sb, the least electronegative central atoms, has the opposite effect : Shared electrons are attracted away from Sb, reducing repulsions between the Sb–I bonds. The consequence is that the effect of the lone pair is greatest in SbI_3, which has the smallest angle.

Atomic size arguments can also be used for these species. Larger outer atoms result in larger angles ; larger central atoms result in smallest angles.
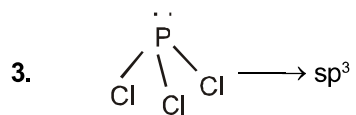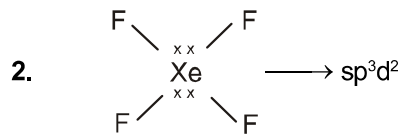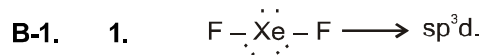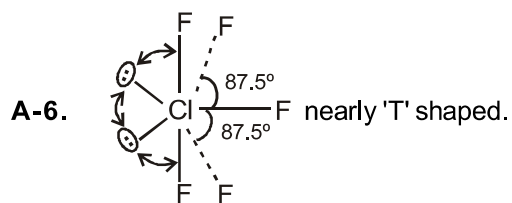
C-2. In Br_3^- there are 5 electron pairs around central bromine atom (3 lone pairs and 2 bond pairs). The hybridisation of bromine is thus sp^3d. To have minimum repulsions between lp-lp and lp-bp it acquires linear shape as shown below.
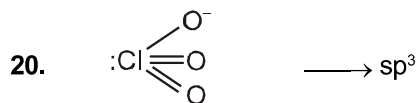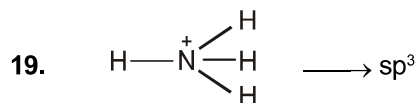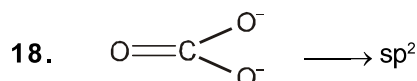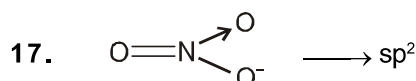

linear
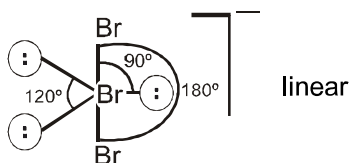
C-3. Order of C–H bond strength is $C_2H_2 > C_2H_4 > C_2H_6$ as % s character decreases in the same order.

C-4.
Highest bond angle	Lowest bond angle
(1) CH_4	SbH_3
(2) CO_2	H_2Te
(3) H_2O	PH_3
(4) ClO_2	Cl_2O
(5) PF_3 (sp^3 hybridisation)	PH_3 (no hybridisation)
(6) BF_3 (sp^2 hybridisation)	NF_3 (sp^3 hybridisation)
(7) NH_3	NF_3
(8) PCl_3	PF_3

C-5. (1) $C \equiv C < C = C < C - C$

(2) $C - F < C - O < C - N$

(3) $HF < H - Cl < H - Br < H-I$

D-1. Structure of P_4O_{10}.

σ bonds = 16
π bonds = 4
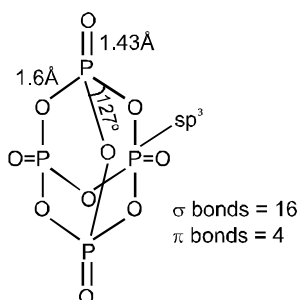
D-2. (a) 4   (b) 3   (c) 3   (d) 6   (e) 4   (f) 0

D-3. **Ans.** 2 (a,d)

(a) P_4O_{10},
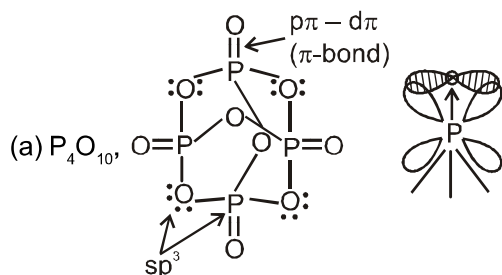
(b) and (c) nitrogen does not have empty d-orbital.

(d) $HClO_3$,
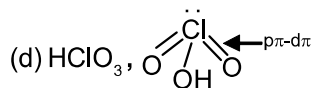
Part - II

A-1.	(C)	**A-2.**	(D)	**A-3.**	(D)	**A-4.**	(A)	**A-5.**	(D)
B-1.	(B)	**B-2.**	(A)	**B-3.**	(A)	**B-4.**	(B)	**B-5.**	(B)
B-6.	(A)	**B-7.**	(D)	**B-8.**	(C)	**B-9.**	(A)	**C-1.**	(D)
C-2.	(B)	**C-3.**	(D)	**C-4.**	(B)	**C-5.**	(C)	**C-6.**	(B)
D-1.	(C)	**D-2.**	(D)	**D-3.**	(A)	**D-4.**	(C)	**D-5.**	(B)
D-6.	(D)	**D-7.**	(A)						

Part - III

1. (A – p , q , r) ; (B – q , r , t) ; (C – s) ; (D – r)   **2.** (A – r) ; (B – q) ; (C – s) ; (D – p, t)

3. (A–r,s); (B–p,q,s) ; (C–q); (D–q,r)

Exercise - 2
Part - I

1.	(D)	2.	(C)	3.	(B)	4.	(C)	5.	(B)
6.	(C)	7.	(A)	8.	(B)	9.	(A)	10.	(D)
11.	(C)	12.	(B)	13.	(B)	14.	(D)	15.	(C)
16.	(C)	17.	(A)	18.	(D)	19.	(A)	20.	(A)
21.	(B)	22.	(A)	23.	(D)						

Part - II

1. 10
2. 5 (BF_3, BCl_3, CO_3^{2-}, SO_3, XeF_4)
3. 6
4. 16
5. 4
6. 10
7. 6

Part - III

1.	(AC)	2.	(CD)	3.	(ABC)	4.	(AC)	5.	(BCD)
6.	(BCD)	7.	(ABD)	8.	(ABC)	9.	(BCD)	10.	(ABC)
11.	(ABC)	12.	(BD)							

Part - IV

1.	(D)	2.	(D)	3.	(C)	4.	(C)	5.	(D)
6.	(D)	7.*	(ABC)							

Exercise - 3
Part - I

1. 0 or 8
2. (D)
3. 4
4.* (ACD)
5. **(5 or 6)**
6. **(4.00)**

Part - II

1.	(2)	2.	(4)	3.	(3)	4.	(3)	5.	(2)
6.	(4)	7.	(3)	8.	(1)	9.	(3)	10.	(4)
11.	(3)	12.	(3)							

Chemical Bonding -3
Exercise - 1
Part - I

A-1. (i) In NH_3 molecule N atom has lone pair in sp^3 hybrid orbital while in PH_3 as suggested by its bond angle ($\approx 92°$) the lone pair must be present in 'S' orbital which is much more contracted than sp^3. Hence PH_3 becomes a poor donor than NH_3.

(ii) CH_3CN has lone pair on sp hybridized nitroge n atom while NH_3 has lone pair on sp^3 hybridized nitrogen atom.

A-2. In BCl_3, octet of Boron is incomplete. In SiF_4, silicon has vacant d-orbitals, by which it can accept electron pair.

A-3. BF_3 molecule being electron deficient gets stabilised through $p\pi–p\pi$ back bonding. where as BH_3 removes its electron deficiency through dimerisation and thus exists as B_2H_6.

A-4. sp^3 hybridised orbital of both aluminium and sp^3 hybridised orbital of carbon.

B-1. In Me$_3$N $\longrightarrow$ $\overset{\overset{\displaystyle F}{|}}{\underset{\underset{\displaystyle F}{|}}{B}}$—F, the electron deficiency of boron is compensated by the lone pair of electron donated

by nitrogen atom. Where as in BF$_3$ it is compensated by back bonding with F atom; back bonding is delocalised thus B – F bond has partial double bond character.

B-2. (Shows pπ – dπ bonding) $\underset{\displaystyle H}{\overset{\displaystyle H}{\underset{H}{>}}}$Si = N = C = O $\quad$ sp

(Cannot show pπ – dπ bonding) $\overset{\displaystyle H}{\underset{\displaystyle H}{H—\overset{|}{\underset{|}{C}}—\ddot{N}}}\overset{C}{\underset{sp^2}{=}}\underset{O}{=}$

B-3. (i) $\underset{O}{\overset{\displaystyle O}{\overset{||}{O=S=O}}}$ $\;$ two pπ - dπ bond and one pπ – pπ bond.

(ii) $\underset{\underset{\displaystyle OH}{|}}{HO{\overset{\overset{\displaystyle O}{||}}{-}}P{-}OH}$ one pπ-dπ bond

(iii) N $\equiv$ N two pπ-pπ bond.

(iv) $H — O — \overset{\overset{\displaystyle O}{||}}{\underset{\underset{\displaystyle O}{||}}{Cl}}{=}O$ three pπ-dπ bond.

C-1. Very strong hydrogen bonding occurs in the alkali metal hydrogen fluorides of formula M[HF$_2$]; there is a linear symmetrical anion having an over all, F–H–F distance of 2.26 Å.

[F – H – – – F]$^-$ $\longleftrightarrow$ [F – – – H – F]$^-$

F$^-$ + HF $\longrightarrow$ [FHF]$^-$; ΔH = – 161 ± 8 kJ mol^{-1}
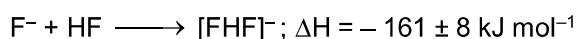
C-2. Deuterium is more electropositive than hydrogen. Therefore, stronger H-bonding is found in D$_2$O than in H$_2$O. D$_2$O is also denser than H$_2$O.

D-1. These compounds contain polar–OH groups which can form H-bonds with water.

D-2. In ethanol, there is H-bonding but in diethyl ether, there is no H-bonding (because O-atom is attached to C-atom) but there exists weak dipole-dipole attraction in diethyl ether.

Part - II

A-1.	(A)	**A-2.**	(A)	**A-3.**	(B)	**A-4.**	(B)	**A-5.**	(D)
A-6.	(B)	**A-7.**	(D)	**B-1.**	(D)	**B-2.**	(A)	**B-3.**	(C)
B-4.	(C)	**B-5.**	(A)	**B-6.**	(B)	**C-1.**	(D)	**C-2.**	(D)
C-3.	(D)	**C-4.**	(D)	**D-1.**	(D)	**D-2.**	(B)	**D-3.**	(D)
D-4.	(C)								

1. (A - q) ; (B – r) ; (C – s,p) ; (D – p, r)
2. (A – p, q, r, s) ; (B – p, q, r, s) ; (C – p, q, r, s) ; (D – r, s)
3. (A – r) ; (B – p, r) ; (C – p, q, r) ; (D – p, r) ; (E – p, r, s)

Exercise - 2
Part - I

1.⋧	(D)	2.	(C)	3.⋧	(B)	4.⋧	(B)	5.	(A)
6.	(D)	7.	(B)	8.	(C)	9.	(B)	10.	(B)
11.⋧	(C)	12.	(A)	13.⋧	(D)				

Part - II

1.	4	2.⋧	20	3.⋧	07	4.	8

Part - III

1.⋧	(BC)	2.	(AC)	3.	(AB)	4.⋧	(ABCD)	5.	(AB)
6.⋧	(ABD)								

Part - IV

1.	(B)	2.	(B)	3.⋧	(D)	4.	(C)	5.	(C)
6.	(C)	7.	(D)	8.	(D)				

Exercise - 3
Part - I

1. **(i)** N(SiMe$_3$)$_3$ is trigonal planar because in it silicon uses its vacant d-orbital for pπ-dπ back bonding with lone pair of electrons of central N-atom and the pπ-dπ bonding is delocalised as given in the structure. So, N(SiMe$_3$)$_3$ with steric number three is trigonal planar.
   (ii) In N(Me$_3$) , there is no such pπ-dπ delocalisation of lone pair of electrons on N atom as carbon does not have vacant d-orbital. So N(Me)$_3$ with steric number four is trigonal pyramidal with a lone pair at the apex.
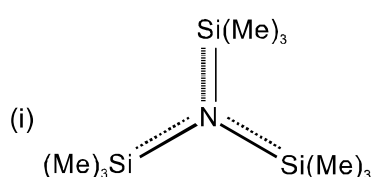
Hence both are not isostructural.
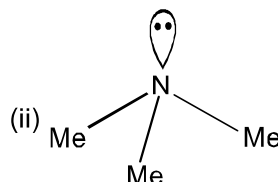
2. **4**

Part - II

1.	(2)	2.	(4)	3.	(4)	4.	(2)	5.	(2, 4)
6.	(4)								

Chemical Bonding - 4
Exercise - 1
Part - I

A-1.⋧ (a) 1   (b) 1/2   (c) 0   (d) 1   (e) 0   (f) 1

A-2. (a) B$_2$ (b) C$_2$ (c) O$_2^{2+}$ (d) O$_2$, (e) F$_2$ (f) N$_2$

A-3. Molecular orbital electronic configuration of O$_2$ is as follows (Z is taken as molecular axis).

$$\sigma_{1s}^2 \; \sigma_{1s}^{*2} \; \sigma_{2s}^2 \; \sigma_{2s}^{*2} \; \sigma_{2p_z}^2 \; \pi_{2p_x}^2 = \pi_{2p_y}^2 \; \pi_{2p_x}^{*1} = \pi_{2p_y}^{*1}$$

Bond order = $\dfrac{10-6}{2}$ = 2.

As it contains two unpaired electrons in bonding π molecular orbitals O$_2$ is paramagnetic.

So, Magnetic moment = $\sqrt{n(n+2)}$ = $\sqrt{2(2+2)}$ = 2.83 B.M.

A-4. 3

B-1. The electronic configuration of O_2 will be:

$$O_2 = \sigma 1s^2\ \sigma^*1s^2\ \sigma 2s^2\ \sigma^*2s^2\ \sigma 2p^2_z\ \pi 2p^2_x = \pi 2p^2_y\ \pi^*2p^1_x = \pi^*2p^1_y$$
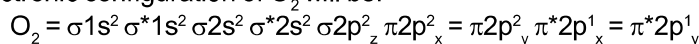
Now   bond order $= \dfrac{N_b - N_a}{2}$

Where, N_b = Number of electrons in bonding orbitals N_a = Number of electrons in antibonding orbitals

bond order $= \dfrac{10-6}{2} = 2$

Similarly electronic configuration of O_2^- (in KO_2) will be

$$\sigma 1s^2\ \sigma^*1s^2\ \sigma 2s^2\ \sigma^*2s^2\ \sigma 2p^2_z\ \pi 2p^2_x = \pi 2p^2_y\ \pi^*2p^2_x = \pi^*2p^1_y$$
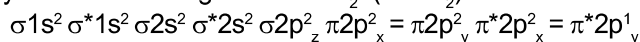
Bond order $= \dfrac{10-7}{2} = \dfrac{3}{2} = 1.5$

In $O_2\ [AsF_4]^-$, O_2 is O_2^+ .
The electronic configuration of O_2^+ will be

$$\sigma 1s^2\ \sigma^*1s^2\ \sigma 2s^2\ \sigma^*2s^2\ \sigma 2p^2_z\ \pi^2p^2_{x'} = \pi 2p^2_y\ \pi^*2p^1_x$$
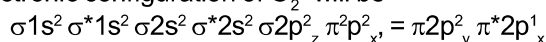
bond order $= \dfrac{10-5}{2} = 2.5$

Hence bond length order will be $O^+_2 < O_2 < O^-_2$ because Bond order $\propto \dfrac{1}{\text{Bond length}}$.

B-2.   Boron (B_2) : B_2 is a good example of the energy level shift caused by the mixing of s and p orbitals. In the absence of mixing, the σ_g (2p) orbital is expected to be lower in energy than the π_u(2p) orbitals and the resulting molecule would be diamagnetic. However, mixing of the σ_g(2s) orbital with the σ_g(2p) orbital lowers the energy of the σ_g(2s) orbital and increases the energy of the σ_g(2p) orbital to a higher level than the π orbitals, giving the order of energies shown below. As a result, the last two electrons are unpaired in the degenerate (having the same energy) π orbitals, and the molecule is paramagnetic.
$(\sigma 1s)^2\ (\sigma^*1s)^2\ (\sigma 2s)^2\ (\sigma^*2s)^2\ (\pi 2p^1_x = \pi 2p^1_y)\ (\sigma p_z)^0$.
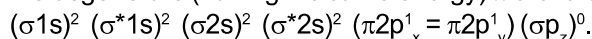
B-3. NO^+ and NO are derivative of N_2 ; so NO^+ bond order = 3 and NO bond order = 2.5 ; B.O. $\propto$ bond strength.

B-4. (ii) & (iv)

B-5. O–O bond length order is ii < i < iii

C-1. Weakest metallic bonding amongst the 3d – series elements $\rightarrow$ no unpaired electrons available for metallic bonding in case of zinc.

C-2. Be should have higher melting point as it contain 2 electrons for metallic bonding where as Li contain only one. Further more, size of Be is smaller than that of Li.

Part - I

A-1.	(C)	**A-2.**	(C)	**A-3.**	(A)	**A-4.**	(A)	**A-5.**	(C)
A-6.	(B)	**A-7.**	(D)	**A-8.**	(D)	**A-9.**	(B)	**A-10.**	(C)
A-11.	(B)	**B-1.**	(A)	**B-2.**	(A)	**B-3.**	(B)	**B-4.**	(A)
B-5.	(A)	**B-6.**	(C)	**B-7.**	(D)	**B-8.**	(D)	**C-1.**	(D)
C-2.	(D)	**C-3.**	(D)						

PART - III

1. (A – r) ; (B – p) ; (C – s) ; (D – q)   **2.** (A – p, r) ; (B – p, q, s) ; (C – q, s) ; (D – p, r)

Exercise - 2
PART - I

1.	(A)	**2.**	(A)	**3.**	(C)	**4.**	(B)	**5.**	(A)
6.	(D)	**7.**	(B)	**8.**	(C)	**9.**	(A)	**10.**	(D)
11.	(D)	**12.**	(D)	**13.**	(A)	**14.**	(D)	**15.**	(D)
16.	(C)	**17.**	(D)	**18.**	(A)	**19.**	(C)		

PART - II

1. ✎ 4
2. ✎ 6
3. 9 (NO_3^-, CO_3^{2-}, F_2, Cl_2, Br_2, O_2^{2-}, O_2^-, Li_2^+, He_2^+)
4. 7

PART - III

| 1.	(ABC)	2.	(ABC)	3.	(BD)	4.	(ABC)
5.	(ABCD)	6.	(AC)	7.	(BC)	8.	(ABD)

PART - IV

1.	(D)	2.	(D)	3.	(B)	4.	(C)	5. ✎	(C)
6.	(D)	7.	(C)	8.	(D)				

EXERCISE - 3

PART - I

1. (A) - p, q, r, t ; (B) - q, r, s, t ; (C) - p, q, r ; (D) - p, q, r, s

2.	(A)	3.	(C)	4.	(C)	5.	(AC)

PART - II

1.	(1, 2)	2.	(3)	3.	(2)	4.	(3)	5.	(2)
6.	(1)	7.	(1)						