

Chemistry Handout For Grade -11

Unit-3 – Chemical Bonding and Atomic Structure

Chemical bond – attractive forces that hold or bind atoms or ions together in the molecule or compound.

- There are three main type of chemical bond (Ionic, Covalent and Metallic).

2009. The force that hold different atoms or ions together are called

- A. Electrical force
- B. Physical force
- C. Chemical bond
- D. atomic force

2008. (Bon→Q9). Which of the following is NOT a form of main chemical bonding?

- A. Covalent Bonding
- B. Metallic Bonding
- C. Ionic Bonding
- D. Hydrogen Bonding

Ionic Bonding (IB) is also called **electrovalent bond**

- Formed between :-
 - i. Metal (low ionization energy) and non metal (high electron affinity)
 - ii. Metal (NH_4^+) and Non metal (radical)

Radical (compound ion) $\rightarrow \text{OH}^-, \text{NO}_3^-, \text{CO}_3^{2-}, \text{NO}_2^-, \text{SO}_4^{2-}, \text{SO}_3^{2-}, \text{PO}_4^{3-}, \text{ClO}_4^-$

- ❖ Compound of one metal and one non metal contain only ionic bond
Example: - NaCl, MgBr₂, CaCl₂, KBr, KI,
- ❖ Compound of one metal and two non metals (radical) contain both ionic bond and covalent bond.
Example: - Na₂CO₃, MgSO₄, CaCO₃, KNO₃, K₃PO₄,
- ❖ Compound NH_4^+ and non metal contain ionic, covalent and dative bond
Example: - NH₄Cl, NH₄Br, NH₄I, (NH₄)₂CO₃, (NH₄)₂SO₄

2010. Which of the following **DOES NOT** contain an ionic bond?

- A. HCl
- B. LiH
- C. NaOH
- D. K₂S

2002. Which of the following compounds **UNLIKELY** to contain ionic bond?

- A. CO
- B. LiCl
- C. NaF
- D. MgBr₂

2002. Four elements A, B, C, and D have atomic number z-1, z, z+1 and z+2 respectively. If z is 9, then the bond between which pair of elements will be ionic?

- A. B & D
- B. A & C
- C. C & D
- D. B & C

A. z-1 = 9-1 = 8 (Oxygen $\rightarrow$ non metal)

B. z = (Fluorine $\rightarrow$ non metal)

C. z + 1 = 9+1 = 10 (Neon $\rightarrow$ non metal)

D. z+2 = 9+2 = 11 (Sodium $\rightarrow$ metal)

$\rightarrow$ Ionic bond is formed by A & D and B & D

2006. (Ionic Q30). Which compound contains both ionic and covalent bond?

- A. Sodium carbonate (Na₂CO₃)
- B. Magnesium bromide (MgBr₂)
- C. Ethanoic acid (CH₃COOH)
- D. Dichloro ethane (CH₂Cl₂)

1999. Which of the following compound contains both ionic and covalent bond?

- A. CaCO₃
- B. PCl₃
- C. MgF₂
- D. CH₂O

Empirical Formula (EF) – is the simplest chemical or molecular formula.

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❖ Step or rule to write empirical formula

- i. Write metal and non metal with their charge respectively

Charge of metal = group of metal = Z_{eff} = valence electron of metal

Charge of non metal = Group of non metal – 8 = Z_{eff}

- ii. Use criss – cross method to write empirical formula.

- iii. Write radical in (Radical)

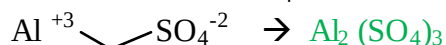
2009. What is the chemical formula of the compound formed by Al (III) & sulfate (SO₄⁻²) ?

A. AlSO₄

B. Al₂SO₄

C. Al₂(SO₄)₃

D. Al(SO₄)₃



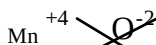
2006. (Simp. Q₇). The simplest formula for a compound containing Mn⁺⁴ & O⁻² is

A. MnO

B. Mn₂O₄

C. MnO₂

D. Mn₄O₂



Mn₂O₄ MnO₂

2007. (Comb. Q₇₂). A compound is formed by the combination of V⁺¹ & X⁻³ is

A. VX

B. V₂X

C. V₃X

D. V₃X₂

2005. Which of the following ionic compounds is formed from the reaction between magnesium and nitrogen?

A. MgN₂

B. Mg₂N₂

C. Mg₃N₂

D. Mg₂N₃

Describing Ionic Bonds

Lewis electron dot symbols → **Lewis dot formula**

- ❖ Representation of valence electrons as dot around symbol of element.

- ❖ Used to shows transfer of electrons between metal and non metal in the formation of ionic bond.

- ❖ Shows formation of ionic bonds by transfer of electrons

Number of dots = Number of valence electrons = Group of element = Z_{eff}

- ❖ One dot represent on electron

Group	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
Period-2	Li.	.Be.	.B.	.C.	:N.	:O:	:F:	:Ne:
Period-3	Na.	.Mg.	.Al.	.Si.	.P:	:S:	:Cl:	:Ar:

Number of ionic bond = Number of transferred electrons by metals to non metals

Q. How many ionic bonds formed by:-

A. NaCl → 1x1 = 1

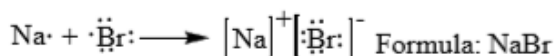
B. MgBr₂ → 2 x 1 = 2

C. Li₃N → 3x1 = 3

D. Al₂(SO₄)₃ → 3x2 = 6

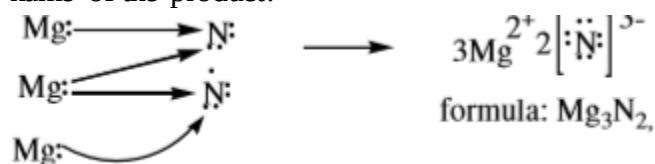
Exercise 3.2.

- 1) Use Lewis electron-dot symbols to depict the formation of sodium and bromide ions from the atoms and determine the formula of the compound.

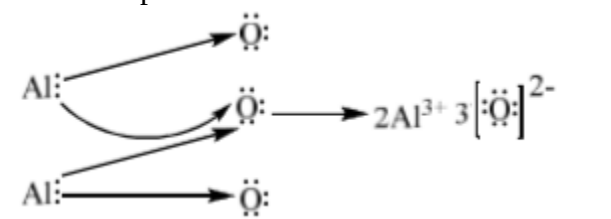


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- 2) Use Lewis electron-dot symbols to show the transfer of electrons from magnesium atoms to nitrogen atoms to form ions with noble gas electron configurations. What is the formula and name of the product?



- 3) Use Lewis electron-dot symbol to show the transfer of electron from aluminium to oxygen atoms to form ions with noble gas electron configurations. What is the formula and name of the product?



Properties of ionic compounds

Ionic compounds

- Ionic compounds have high melting point, boiling pint and density.
- Ionic compounds are solid at room temperature
- Ionic compounds are composed of ions (cation and anion)
- Ionic compounds are electrolyte (conductor) in liquid state, molten state, fused state and aqueous solution
- Ionic compounds are non conductors/electrolyte in solid state.
- Ionic compounds are non volatile
- Ionic compounds are soluble in polar or inorganic solvents.
- Ionic compounds are soluble in non polar solvents or organic solvents
- Ionic compounds are hard and brittle.

Factors affecting formation of ionic compounds

- Ionization energy (IE)
 - Electron affinity (EA)
 - Lattice energy (U)
- ✓ Formation of ionic bond is **favoured** or affected **low ionization energy**, **high electron affinity** and **high lattice energy**

Lattice energy (U)

- Is coulombic **attractive force between ions**
- Is **energy required to break one molecule of ionic solid into isolated gaseous ions.**
- **Indicate strength of ionic bond** (**U increases**→ **strength of ionic bond increases**)
- Plays crucial role in the ionic bond formation
- **Determines many properties of ionic compounds.**

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- As **lattice energy increases** →
 - Strength (hardness) of ionic bond increases
 - **Melting point of ionic compound increases**
 - **Boiling point of ionic compound increases**
 - **Ionic charge increases**
 - Rate of reaction increases
 - **Ionic size or radius decreases**
 - **Solubility of ionic compounds in water decreases**

Identify :- Lattice **energy of compounds of the same group by ionic size**
Lattice energy of compounds of different group by ionic charge.

Activity 3.6

Q₁. In general, how does the lattice energy of an ionic compound depend on the charges and size of ions?

The lattice is directly proportional to charges of ions and inversely proportional to radius of ions.

Q₂. For each pair, choose the compound with the higher lattice energy and explain.

- A. LiCl or CsCl → **LiCl because Li is much smaller than Cs in size.**
B. NaF or MgO → **MgO the charges on Mg and O are twice that of Na and F.**
C. BaS or CsI → **BaS the charges on Ba and S are twice that of CS and I**
D. Arrange the following ionic compounds in increasing order of lattice energy, melting point, boiling point and solubility.
a. NaCl, LiCl, KCl, RbCl

Ans → **RbCl < KCl < NaCl < LiCl** (increasing order of melting point, boiling point and lattice energy) and **LiCl, NaCl, KCl and RbCl** increasing order of solubility in water.

2007. Which of the ionic compounds has the greatest lattice energy?

- A. **LiF** B. LiCl C. LiBr D. LiI

As size of metal and nonmetal increases → lattice energy, melting and boiling point decreases

1999. Which of the ionic compounds has the smallest lattice energy?

- A. LiF B. NaF C. **CsI** D. LiI

2002. Which of the compounds would be expected to have the highest melting point?

- A. **BaF₂** B. BaCl₂ C. BaBr₂ D. BaI₂

Lattice energy can be calculated using Hess's law through Born Haber cycle.

$$\Delta H_f^0 = \Delta H_{\text{Sub}}^0 + \Delta H_{\text{IE}}^0 + \Delta H_{\text{diss}}^0 + \Delta H_{\text{EA}}^0 + U$$

ΔH_f^0 = Standard Enthalpy of formation (KJ/mole)

ΔH_{Sub}^0 = Standard Enthalpy of sublimation (KJ/mole)

ΔH_{IE}^0 = Standard Enthalpy of ionization energy (KJ/mole)

ΔH_{diss}^0 = Standard Enthalpy of dissociation (KJ/mole)

ΔH_{EA}^0 = Standard Enthalpy of electro affinity (KJ/mole) and

U is lattice energy (KJ/mole)

Born –Haber cycle is a:-

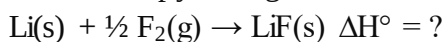
❖ Series change in a closed cycle in which net change in energy must be zero.

❖ Has five series steps

- Sublimation of metals (atomization) → converts solid metals to gas metals.
- Ionization of metals → converts gas metals to gas metal ions
- Dissociation of non metals → converts molecule of non metal to atoms.
- Formation of anion from non metal atoms
- Formation of ionic crystal from metal ion and non metal ions (combine cation and anion

Exercise 3.3

- 4) For lithium, the enthalpy of sublimation is $+161\text{kJmol}^{-1}$ and the first ionization energy is $+520\text{kJmol}^{-1}$. The dissociation energy of fluorine is $+154\text{kJ mol}^{-1}$ and the electron affinity of fluorine is -328kJmol^{-1} . The lattice energy of LiF is -1047kJ mol^{-1} . Calculate the overall enthalpy change for the reaction?



$$\Delta H_f^\circ = \Delta H_{\text{Sub}}^\circ + \Delta H_{\text{IE}}^\circ + \Delta H_{\text{diss}}^\circ + \Delta H_{\text{EA}}^\circ + U$$

$$\Delta H_f^\circ = -1047\text{kJ/mol} + \left(161 + 520 + \frac{154}{2} - 328 \right) \text{kJ/mol} = -617\text{kJ}$$

- 5) The enthalpy of formation of cesium chloride is $\text{Cs(s)} + \frac{1}{2}\text{Cl}_2(\text{g}) \rightarrow \text{CsCl(s)} \quad \Delta H^\circ = -44.28\text{kJmol}^{-1}$. The enthalpy of sublimation of cesium is $\text{Cs(s)} \rightarrow \text{Cs(g)} \quad \Delta H^\circ = +77.6\text{kJmol}^{-1}$. Use these data, with other data from other sources, to calculate the lattice energy of CsCl(s)

Solution:-



$$\Delta H_{\text{Sub}}^\circ = 77.6\text{KJ/mole}, \quad \Delta H_{\text{EACl}}^\circ = -348.7\text{KJ/mole}, \quad \Delta H_{\text{diss}}^\circ = 243.4/2 = 121.7\text{KJ/mole}$$

$$U_{\text{CsCl}} = \Delta H_f^\circ - (\Delta H_{\text{Sub}}^\circ + \Delta H_{\text{IE}}^\circ + \Delta H_{\text{diss}}^\circ + \Delta H_{\text{EA}}^\circ)$$

$$= -44.2\text{KJ} - (77.6 + 774.6 = 121.7 - 348.7\text{KJ})$$

$$= -44.2\text{KJ} - (973.42 - 348.7\text{KJ}) = -44.2\text{KJ} - 626.72\text{KJ} = -669\text{kJ}$$

- 6) Using the following data: Enthalpy of sublimation of Ca = $+178.2\text{kJ mol}^{-1}$, Enthalpy of dissociation of Cl_2 = $+243.4\text{kJ mol}^{-1}$, Enthalpy of formation of CaCl_2 = -795.8kJ mol^{-1} First and second Ionization energies for Ca are $+590\text{kJ mol}^{-1}$ and $+1145\text{kJ mol}^{-1}$ respectively. The electron affinity of Cl = -348.7kJ mol^{-1} Determine the lattice energy of CaCl_2

$$\text{Solution: } U_{\text{CsCl}} = \Delta H_f^\circ - (\Delta H_{\text{Sub}}^\circ + \Delta H_{\text{IE}}^\circ + \Delta H_{\text{diss}}^\circ + \Delta H_{\text{EA}}^\circ) \\ -795.8 - [178.2 + 243.4 + 590 + 1145 + (-348.7 \times 2)] = -2255\text{kJ}$$

- Write the formulas and names of the compounds formed from the following ionic interactions: (use periodic table)
 - The 2^+ ion and 1^- ion are both iso electronic with the atoms of a chemically un reactive period 4 element.
 - Anion (negative ion)** is iso electronic with noble gas atom at ends of its period.
Example: F^- , O^{2-} , N^{3-} ...

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- **Cation (positive ion)** is iso electronic with noble gas atom at preceding of its period. Example: Na^+ , Mg^{+2} , Al^{+3} ...
- **Charge** indicates group of elements.
- **+2 ion** is group IIA element $\rightarrow$ iso electronic with **period 4 noble gas** which found in period 5 (**metal of group IIA and Period 5**) $\rightarrow \text{BeMgCaSr}$
- **-1 ion** group is VIIA $\rightarrow$ iso electronic with **period 4 noble gas** which found in period 4 (non **metal of group VIIA and Period 4**) $\rightarrow \text{FCIBrIAt}$

Ans $\rightarrow \text{SrBr}_2$ (Strontium Bromide)

- b. The 2^+ ion and the 2^- ion are both iso electronic with the period 3 noble gas.
- $+2$ ion iso electronic with period $\rightarrow$ metal of group IIA and period 4 (Ca)
 - -2 ion iso electronic with period 3 $\rightarrow$ non metal of group VIA and period 3 (S)

Ans $\rightarrow \text{CaS}$ (calcium sulphide)

- c. The ions formed are the largest and smallest ionizable atoms in period 2.

Largest period 2 cation is Li^+ and smallest period 2 is F^-

Ans $\rightarrow \text{LiF}$ (lithium fluoride)

2. In each of the following ionic compounds identify the main group to which X belongs:

a. XF_2

c. X_2O_3

b. MgX

d. Na_2X

a. Group IIA

b. Group VIA

c. Group IIIA

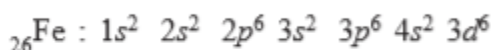
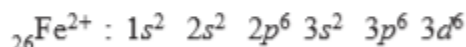
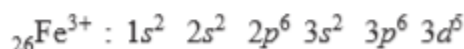
d. Group VIA

Exceptions to octet rule of ionic compounds

- i. Ionic compounds of Be^{+2} , Li^+ , H, transition metals and post transition metals (Ga, In, Ge, Pb, Sn) don't obey/violate/ octet rule.
 - ii. The heavier post-transition elements like Pb and Sn lose the p electrons or both the p and s electrons from the valence shell.
- ✓ Example :- LiH , Li_2O , LiF , BeCl_2 , CuO , FeCl_3 , ZnO , ZnCl_2 , PbO , SnO , SnO_2 , PbO_2

Home work: Write electron configuration for Fe^{+2} , Zn^{+3} , Cu^{+2} , Ga^{+3} which obey pseudo noble gas configuration.

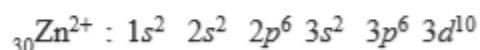
Electron Configurations of Iron:



- ❖ Fe^{2+} is a well-known stable ion with a valence shell electron configuration $3s^2 3p^6 3d^6$ which is **not iso electronic** with any of the noble gases.

Electron configurations of zinc:

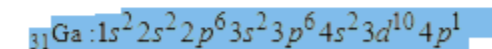
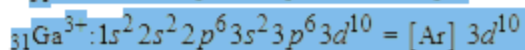
Also,



is not isoelectronic with any of the noble gases.

Electron configurations of gallium:

The post-transition element gallium (Ga) loses electrons first from the 4p orbital and then from the 4s orbital to form a Ga^{3+} ion as



- ✓ Electron configurations of Zn^{2+} and Ga^{3+} ions have completely-filled outer subshells and a noble gas core.
- ✓ Their valence electron configuration can be generally represented as $ns^2 np^6 nd^{10}$.

Covalent (electron pair sharing) bond

- ❖ Is formed by sharing pairs of electrons between non-metals (metalloid and non metals).

Example: - H_2O , HCl , CH_4 , PCl_3 , NH_3 , SiO_2

Bonding pairs (BP)	Lone pairs(LP)
<ul style="list-style-type: none">▪ Pair of electrons remain or exist between two atoms▪ Bonding electrons▪ electrons involve in bond formation▪ Shared pair of electrons between two atoms▪ Written (-) or : between atoms	<ul style="list-style-type: none">▪ Pair of electrons remain on one atom/element/▪ Non- bonding electrons▪ Un- shared pair of electrons▪ Electrons that do not involve in bond formation▪ Written: on one element.

Types of Covalent bonding

1) Single Bond(SB)

- ❖ Covalent bond formed by sharing one pair (two) electrons
- ❖ One pair bond

Example: - F_2 , Cl_2 , Br_2 , I_2 , H_2 , H_2O , HCl , CH_4 , NH_3 , Alkane

2) Double Bond (DB)

- ❖ Covalent bond formed by sharing two pair (four) electrons
- ❖ Two pair bond

Example: Alkenes, CO_2 , C_2H_4 , O_2 , SO_2 , SO_3 , O_3

3) Triple Bond(TB)

- ❖ Is covalent bond formed by sharing three pair (six) electrons
- ❖ Three pair bond

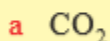
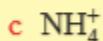
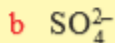
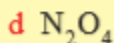
Example: Alkynes, CO , C_2H_2 , N_2 , HCN

Rules to write the Lewis formula/structure/ for a covalent compound/molecule/

1. **Determine the total number of valence electrons** and number of pairs (BP and LP).
 - ✓ The total number of electrons for a **molecule** is the sum of the valence electrons for each atom.
 - ✓ For a **polyatomic anion**, which has one or more extra electrons, add one electron for each unit of negative charge.
 - ✓ For a **polyatomic cation**, which is missing one or more electrons, subtract one electron for each unit of positive charge.
$$\text{Number of pair} = \frac{\text{Total valance electrons}}{2} = \frac{\text{BP} + \text{LP}}{2}$$
2. Predict /decide/ **central atom** or element
 - **Central atom** is most electropositive atom or fewer atom
 - **Terminal or surrounding atoms** – most electron negative atoms (F)
 - Hydrogen and F are always a terminal atom
 - O and halogen usually surrounding atoms
3. **Place electron pairs between central and terminal atoms**
4. **Assign any remaining/extra/ electrons as lone-pairs around the central atom**
5. **If octet rule of central atom is not complete.**
 - ✓ Make multiple bond (DB and TB) between central atom and surrounding atoms by removing lone pair from surrounding atoms
 - ✓ Terminal oxygen is **not** shares its electron to form single bond with central atom
 - ✓ Terminal oxygen is shares its electron to form multiple bond with central atom

Exercise 3.4

Determine the total number of valence electrons for the following species:



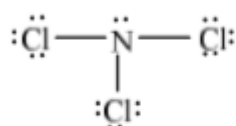
Answers

- | | |
|--|--|
| a. CO_2 has $(1 \times 4) + (2 \times 6) = 16$ | b. SO_4^{2-} has $(1 \times 6) + (4 \times 6) + 2 = 32$ |
| c. NH_4^+ has $(1 \times 5) + (4 \times 1) - 1 = 8$ | d. N_2O_4 has $(2 \times 5) + (4 \times 6) = 34$ |

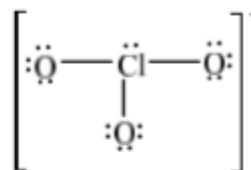
Exercise 3.5

- Write a plausible Lewis structure of:
 - nitrogen trichloride, NCl_3
 - chlorate ion, ClO_3^-
 - phosphonium ion, PH_4^+
 - phosgene, COCl_2
- Draw a Lewis structure for CO_3^{2-} , SF_4 and HCOOH (formic acid).
- Which of the following atoms cannot serve as a central atom in a Lewis structure O, He, F, H, P? Explain.
- Write a plausible Lewis structure for carbonyl sulphide, COS .

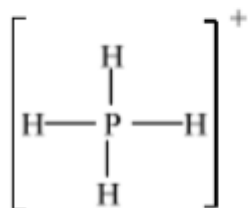
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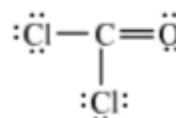
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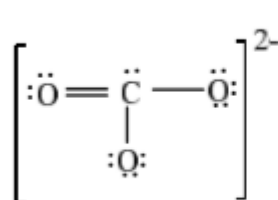
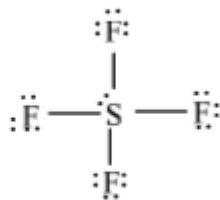
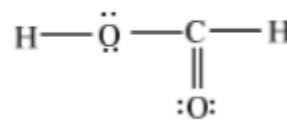
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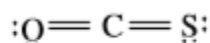
d.



2.



3. F and H



4.

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2004. What is the total number of valence electrons in BrO_3^- ?

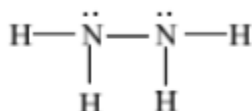
- A. 20 **B. 26** C. 32 D. 36

▪ Total number of valence electrons in $\text{BrO}_3^- = (1 \times 7) + (3 \times 6) + 1 = 26$

2010. When a student draws a plausible Lewis Structure for hydrazine (N_2H_4) molecules, how many lone pair of electrons are available? Review Q₃₁

- A. 1 **B. 2** C. 3 D. 0

Solution



2009. Which of the following Lewis structure for oxygen atom?

- A. Three lone pairs of valence electrons
B. One lone pair of valence electrons and one bonding sites
C. Two lone pairs of valence electrons and one bonding site
D. Two lone pairs of valence electrons and two bonding site
✓ O atom has two lone pairs and two bonding sites

Number of bonding sites = number of single electrons

2009. How many electrons are shared between two atoms in a double bond?

- A. 2 B. 3 C. 4 D. 6

- ✓ **Single bond** contains **two** shared electrons
✓ **Double bond** contains **four** shared electrons
✓ **Triple bond** contains **six** shared electrons

2008. Which of the following statements is **NOT** true about covalent bonding?

- A. Covalent bonds are least likely to be formed between atoms of the same element.
Example: H_2 , Cl_2 , F_2 , Br_2 , I_2 , O_2 , N_2
B. Covalent bonds are least likely to be formed between atoms of different elements on the right side of the periodic table.
Example: CO_2 , SO_2 , F_2O
C. Covalent bonds are least likely to be formed between an element in Group I and an element in Group VII.
Form ionic bond
D. Covalent bonds are least likely to be formed by head o the group elements with high ionization energies.

Covalent bond is formed by nitrogen, oxygen, halogen ...

2008. How many bonding pairs and lone pairs, respectively does the ion ICl_4^- have?

- A. 3, 2 **B. 4, 2** C. 5, 1 D. 4, 1

2007. How many unpaired electrons are there in the Lewis structure of N^{3-} ?

✓ All electrons are paired

2005. The total number of electrons participating in the bond formation of carbonate anion CO_3^{2-} in the molecule of carbonic acid is:

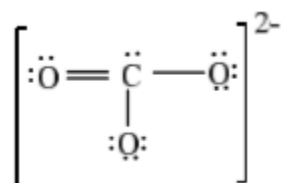
Chemistry Handout For Grade -11

A. 16

B. 10

C. 8

D. 5



- Total valence electrons = $1 \times 4 + 3 \times 6 + 2 = 24$ electrons
- Total pair of electrons = $24/2 = 12$
- Central atom is C
- Bond pairs = 4, shared electrons = 8, lone pairs = 8 and unshared electrons = 16

2004. Which term describes the unit make up to compounds with covalent bond?

A. Ions

B. Acids

C. Salts

D. Molecule

2003. Which set contains only covalently bonded molecules?

A. BCl_3 , SiCl_4 , PCl_3

C. I_2 , H_2S , NaI

B. NH_4Br , N_2H_4 , HBr

D. Al , O_3 , AS_4

Home work. 2001. Which of the following Lewis structure is correct for tri iodide (I_3^-)?

2001. A covalent bond is **unlikely** exist in which of the following substance?

A. H_2

B. SeH_2

C. SiF_4

D. CaO

2000. What does the correct Lewis structure for the molecule CCl_4 molecule?

A. 5- bonds

C. 24 – unshared electrons

B. no unshared electrons

D. 10 –electrons

- ✓ Lewis structure of CCl_4 has 4BP (8 shared electrons) and 12 LP (24unshared electrons)

1995. How many pairs of unshared (non-bonding) valence electrons are on the central atom of SF_4 molecule?

A. 0

B. 2

C. 1

D. 3

Coordinate Covalent Bond

- ❖ Also called dative or donor- acceptor bond
- ❖ Formed by sharing of electrons pair donated by one atom
- ❖ Formed between Lewis acid(acceptor) and Lewis base(donar)
- ❖ Formed by Oxygen and Ammonia
- ❖ One atom can “**donate**” two of its electrons to provide the shared pair between itself and an “**acceptor**” atom.
- ❖ **Donor Atom** – provide both electrons
 - ✓ Must contain at least one lone pair on the central atom (example: NH_3)
- ❖ **Acceptor atom**- must contain vacant orbital to accommodate donated electron pairs
 - Example:** Oxygen atom
 - Example: O_3 , H_3O^+ , NH_4^+ , HNO_3 , NH_3BF_3 , POCl_3 , SO_2 , SO_3 , NO_2^- , NO_3^- , CO_3^{2-} and CO contains Coordinate Covalent Bond.

2012. Given the reaction: $\text{NH}_3 + \text{H}^+ \longrightarrow \text{NH}_4^+$ The bond formed between NH_3 and H^+ is

Chemistry Handout For Grade -11

A. Ionic bond B. Covalent Bond C. Hydrogen bonding D. **Coordinate covalent bond**
2000. Which of the following molecules does **not** have coordinate covalent bond?

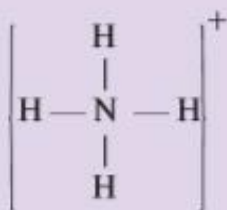
A. H_3O^+ B. NH_4^+ C. NH_3BF_3 D. CO_2

1999. What type of bond formed when a **Lewis acid reacts with Lewis base**?

A. Ionic bond B. Covalent Bond C. Hydrogen bonding D. **Coordinate covalent bond**

Activity 3.9. Discuss the following questions:

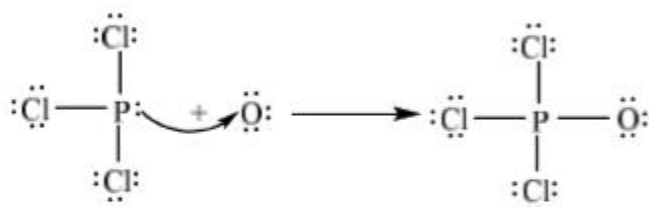
1. Why (a) the ammonia molecule, (b) the oxygen atom can participate readily in coordinate covalent bonding.
2. Use Lewis structures to show the formation of coordinate covalent linkage between (a) phosphorus trichloride and oxygen (b) boron trifluoride and ammonia.
3. In the following structure of ammonium ion, can you identify the coordinate covalent bond?



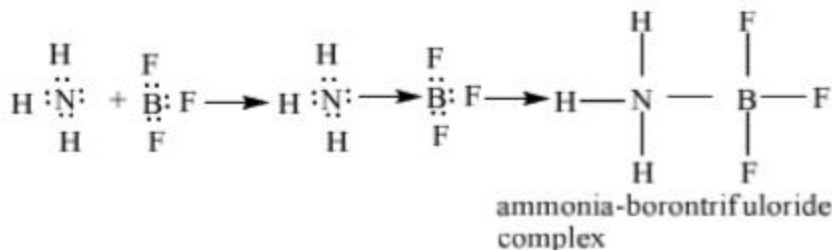
Answer:

1. a. The nitrogen atom in the ammonia molecule has a lone pair of electrons, and it can provide both electrons to an electron-deficient atom or group of atoms.
- b. The oxygen atom can accept an electron pair from donor groups. This is the reason that the ammonia molecule and the oxygen atom participate readily in a coordinate covalent bond.

2. a.



- b. Sometimes, two or more stable molecules combine to form a molecular complex. In a molecular complex, the constituent molecules are held together by a 'coordinate bond'. One typical example involves the molecules of NH_3 and BF_3 . The electron-dot structures of these molecules are:



The nitrogen atom has a complete octet around it, but the boron atom has only six electrons around it. The nitrogen atom therefore donates its lone pair of electrons to boron so that its atoms also acquire the octet. This one-sided sharing between N and B atoms gives rise to a coordinate bond. Coordinate bonds are involved in the formation of transition metal complexes known as coordination compounds.

3. All four hydrogen atoms and all four N-H bonds in the ammonium ion are found by experiment to be equivalent. Therefore, the one which is a coordinate covalent bond can't be identified.

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Resonance Structure is:

- ❖ Also called **contributing structures**
- ❖ Different Lewis structures used to represent the single molecule, compound or ion
- ❖ Cannot be described by single Lewis structure.
- ❖ Differ only in the distribution/position/ or arrangement of electrons
- ❖ Exist in the compound that contain multiple bond
- ❖ **Does not exist** in the molecule /compound/ that contains **only single bond**.
- ❖ Linked each other by double-headed arrows ($\longleftrightarrow$)
- ❖ **Resonance** is the appearance of delocalized electrons within certain molecules or polyatomic ions
- ❖ **Delocalized'** refers to an electron which is not 'attached' to a particular atom.

2010. Number of resonance structure for CO_3^{2-} is:

- A. 9 B. 6 C. 3 D. 2

Number of resonance structure = Number of identical surrounding atoms = 3

2006. (DeliQ68). Which of the following molecules or ions will exhibit delocalized bonding?

NO_2^- , NH_4^+ , N_3^-

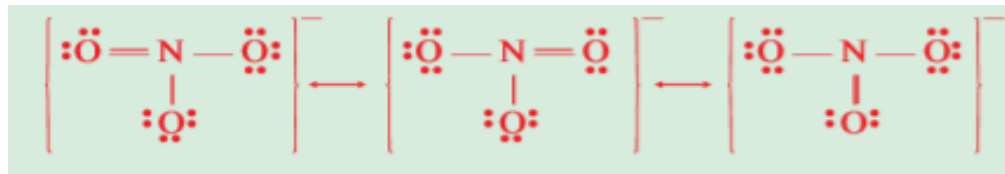
- A. NO_2^- and N_3^- B. NH_4^+ and N_3^- C. NO_2^- D. NO_2^- and NH_4^+

✓ NO_2^- and N_3^- are contain multiple bond (exhibit delocalized bond)

✓ NH_4^+ is contains single bond (does not exhibit delocalized bond)

Example1: Write three equivalent Lewis structures for the nitrate ion, NO_3^- .

Describe its resonance hybrid structure.

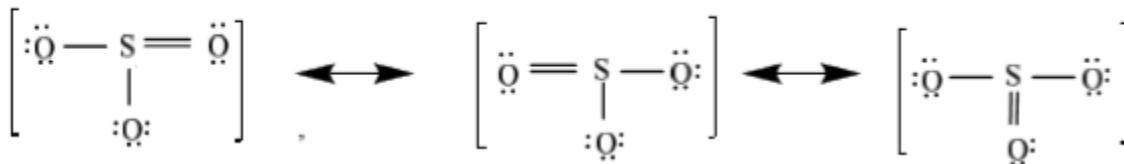


Exercise 3.6

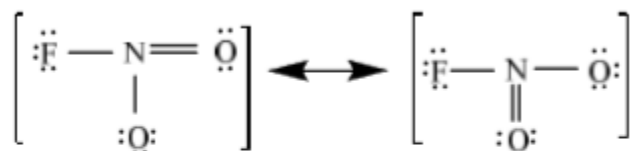
1. Write three equivalent structures for the SO_3 molecule that obey the octet rule.
2. Draw Lewis structures of all the important resonance forms of each of the following:

a NO_2F (N central) b HNO_3 c NO_2^-

1.



2. a.



b.



c.



11. Which of the following species cannot be adequately described by a single Lewis structure?

a OH⁻

c NH₄⁺

b C₂H₂

d HCO₃⁻

Answer: D

Exceptions to the Octet rule in Covalent Bonding

- Molecules or compounds that do not obey octet(violate) octet rule
- There are **three groups of molecules** that are **exceptions to the octet rule**

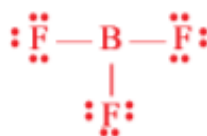
A. **Less than octet** (central atom is deficient of electrons):

- Molecules whose central atoms have fewer than eight electrons (**Below octet**).
- Incomplete octet rule.
- Molecules containing central atoms of Group **IIA and IIIA**.
- Compounds of Be, B, Al

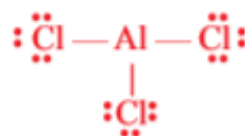
Example: BeCl₂, BF₃ and AlCl₃



4 electrons around Be



6 electrons around B



6 electrons around Al

B. **More than octet** (central atom has excess of electrons):

- Molecules whose central atoms have more than eight electrons (**Expanded octet**).
- Molecules containing central atoms from periods 3, 4, 5, and 6.

Example: PF₅, SF₆, XeF₄, SF₄, BrF₅, I₃⁻, XeF₂, PF₆⁻ are obey expanded octet rule



10 electrons around P



12 electrons around S

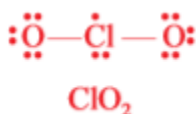


12 electrons around Xe

C. Molecules containing an odd number of electrons:

- Molecule or compound that contains odd or unpaired electrons.

Examples are ClO_2 , NO and NO_2 having 19, 11 and 17 valence electrons respectively.



Exercise 3.7

1. Suggest the Lewis-like structure for NO and NO_2 molecules.
2. The following species do not obey the octet rule. Draw a Lewis structure for each one and state the type of octet rule exception:

a BH_3

b AsF_4^-

c SeCl_4

d PF_6^-

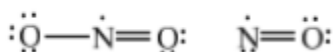
e ClO_3

f BrF_3

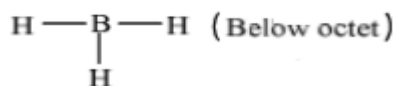
g BeF_2

h XeF_2

1.

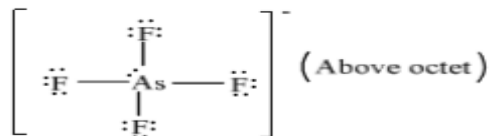
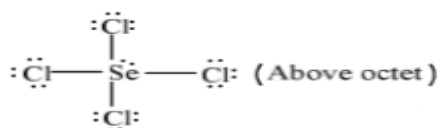


2a.

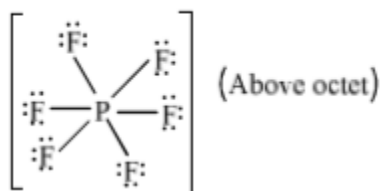


c.

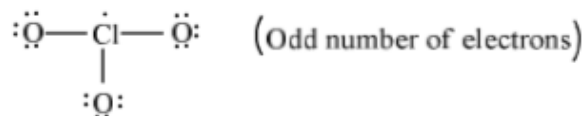
b.



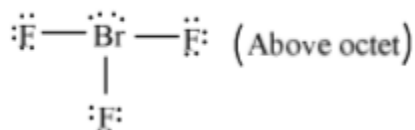
d.



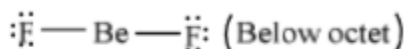
e.



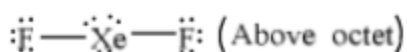
f.



g.

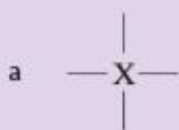
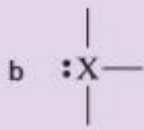
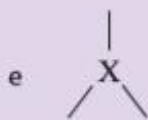
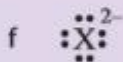


h.



Activity 3.11. Discuss the following questions.

1. In which pattern(s) does X obey the octet rule?



2. What requirements are needed for an atom to expand its valence shell? Which of the following atoms can expand its valence shell; F, S, H, Al?

3. Identify the atoms in PCl_5 , PCl_3 , OF_2 , BF_3 and SiF_4 which do not obey octet rule.

Answer:

1. a, b, and d **obey the octet rule**.

2. Expanded octet occurs in any system that has an atom with more than four electron pairs attached to it.

Example: SF_6 , AsF_5 , XeF_6 , and PF_5

3. Phosphorus in PCl_5 and boron in BF_3 **do not obey the octet rule**

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2007. Which of the following does **NOT** obey octet rule?

- A. CS₂ B. PBr₃ C. IBr D. BrF₅

2004. Which of the following molecules has Lewis structure that does not obey the octet rule?

- A. NO B. CS₂ C. PF₃ D. HCN

Dipole Moment (μ)

- ❖ The dipole moment (μ) is defined as the product of the magnitude of the charge (δ) at either end of the dipole multiplied by the distance (d) that separates the charge.

$$\mu = \delta \times d$$

- ❖ The SI-unit of dipole moment is coulomb-metre (C.m).
- ❖ Dipole moments are often expressed in the non-SI unit debye (D)

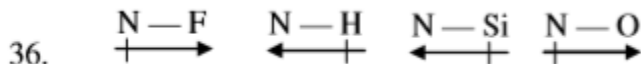
$$1D = 3.33564 \times 10^{-30} \text{ Cm.}$$

- ❖ Measure of polarity of a polar covalent bond.
- ❖ For a diatomic molecule
 - ✓ The bond moment is the dipole moment.
 - ✓ Dipole moment is a vector sum of the bond moments in a molecule.
- Example HCl
- ❖ The dipole moment of a polyatomic molecule (three or more atoms) depends on the geometry of the molecule.
- ❖ If the bond moments are equal in magnitude but opposite in direction, then they will cancel each other and the resultant dipole moment will be zero, provided that the vector sum of the bond moments is zero.
- ❖ Dipole (two poles)
 - ✓ Must be directed from most electro positive atom to the most electro negative atom.
 - ✓ Indicted by cross linked arrow($\longleftrightarrow$)

36. Show the direction of the bond moment (dipole moment) in each of the following bonds:



Answer:



- ❖ Delta plus (δ^+) is written on the partial positive atom and delta minus (δ^-) on the partially negative atom. Example : $\text{H}^{\delta+} - \text{Cl}^{\delta-}$

25. Which atom (if any), in each of the following bonds would you expect to carry a partial negative charge?

- | | | | |
|---|-------|---|-------|
| a | H - N | d | C - F |
| b | S - S | e | B - O |
| c | O - P | f | I - F |

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Answer:

25. a. N
d. F

b. neither
e. O

c. O
f. F

Polarity of (bond & molecule)

I. Polarity of Bond

Covalent bond	
Non-Polar bond	Polar Bond
✓ Covalent bond formed between like/identical / atoms or elements Example: Homo/pure/ molecule (O ₂ , N ₂ , H ₂ , F ₂ , Cl ₂ , Br ₂ , I ₂ , O ₃ , P ₄ , S ₈)	✓ Covalent bond formed between different atoms or elements Example: hetero molecule (CO ₂ , CS ₂ , CCl ₄ , H ₂ O, NH ₃ , AB, XY)

2005. Which of the following molecules represents a non polar covalent bond?

A. B-Cl B. C-Cl C. Cl-Cl D. Mg-Br

2003. Which bond between the following molecules is most polar?

A. S-F B. S-Cl C. S-Cl D. S-I

1995. Which set is arranged in order of increasing polarity?

HI < HBr < HCl < HF

As electro negativity difference increase = bond polarity increase

II. Polarity of molecule or compound

❖ Polar bonds do not necessarily lead to polar molecules

Molecule or compound	
Non-Polar molecule or compound	Polar molecule or compound
i. Opposite bond are identical bond	i. Opposite bond are different bond
ii. Central atom surrounded by identical atom without lone on it	ii. Central atom surrounded by different atom and central atom contain lone pair
iii. Form trans geometry	iii. Form Cis-geometry
iv. $\mu = 0$	iv. $\mu \neq 0$

Polarity of molecule or compound depends on the:

- Polarity of bond
- Geometry or shape of molecule (Trans and cis geometry)
- Electro negativity difference

2007/2005 Which of the following molecules has a dipole moment?

A. XeF₄ B. H₂S C. SO₃ D. CH₄

- ✓ A. Opposite bonds are identical and cancel each other (non-polar molecule)
- ✓ B. has lone pair on the central atom (polar molecule)
- ✓ C and D haven't lone pair on the central atom (non polar molecule)

1996/1998. Which of the following molecule is no-polar?

A. SO₂ B. CS₂ C. CHCl₃ D. SF₄

- ✓ A and D have lone pair on the central atom (polar molecule)
- ✓ C. opposite bonds are different (polar molecule)

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✓ B hasn't lone pair on the central atom (non polar molecule)

8. Carbon and chlorine form a series of compounds: CH_3Cl , CH_2Cl_2 , CHCl_3 , CCl_4 . Which of these will be polar:

- a CCl_4 only
- b CH_3Cl and CHCl_3 only
- c CH_3Cl , CH_2Cl_2 and CHCl_3 only
- d CH_3Cl , CH_2Cl_2 , CHCl_3 and CCl_4

8B.

24. Classify the given molecules as polar or non-polar:

- | | |
|------------------|------------------------|
| a SiF_4 | d H_2S |
| b NF_3 | e CCl_4 |
| c PCl_3 | f CO_2 |

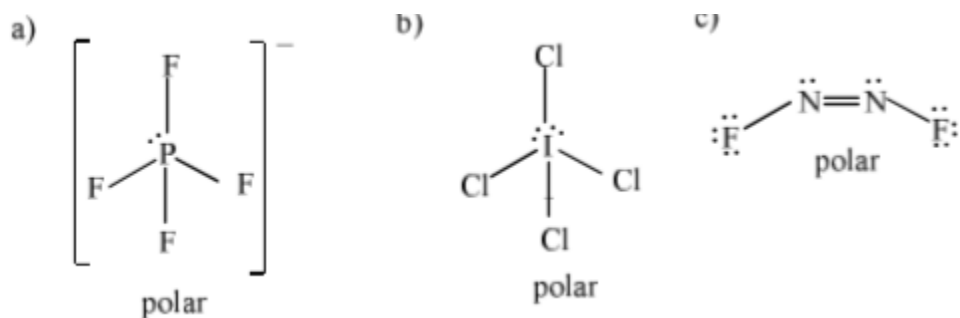
Answer:

- | | | |
|-----------------|-------------|--------------|
| 24. a. nonpolar | b. polar | c. polar |
| d. polar | e. nonpolar | f. non polar |

26. Draw the Lewis structure for the given molecules and state whether the molecule is polar or non-polar:

- | | |
|---|--------------------|
| a PF_4^- | b ICl_4^- |
| c N_2F_2 (2 forms possible) | |

Answer:



Exercise 3.8

Both CO_2 and BCl_3 have zero dipole moments, but the $\text{C}=\text{O}$ and $\text{B}-\text{Cl}$ bond moments are not zero. Explain.

❖ Bond moment in both CO_2 and BCl_3 can be predicted from bond moments cancel to give a non-polar molecule to produce a zero resultant dipole moment.

Activity 3.12.

Collect the following substances. Divide them in two groups. Group A consists of:

- a) Naphthalene
- b) Sodium chloride
- c) Iodine
- d) Wax and

Group B consists of the following solvents

- a) Carbon tetrachloride
- b) Water
- c) Methylbenzene

Each solid in group A dissolves best in at least one solvent in group B. Point out the solid and the solvent it dissolves in and give reasons for your choice.

The answers suggested are as follows:

Solid substance	Solvent
Naphthalene	Carbon tetrachloride
Sodium chloride	Water
Iodine	Methyl benzene
Wax	Methyl benzene

Note: Naphthalene is insoluble in water because it is a fairly non-polar molecule (it is a hydrocarbon, and most hydrocarbons are non-polar) and water is a polar solvent. Wax is soluble in aromatic solvents, chloroform, esters, ethers and acetone. Iodine is more soluble in toluene than water. *Like dissolves like!*

Properties of Covalent Compounds or molecules

- ❖ They are found in gaseous state at room temperature.
- ❖ They are liquids or gases at room temperature.
Liquid — H_2O , Br_2 , Gas — CO_2 , H_2 , Cl_2 , NH_3 .
- ❖ Some covalent molecules like iodine are solids at room temperature.
- ❖ Covalent have low melting and boiling points.
- ❖ They are volatile.
- ❖ Generally they have low melting points and boiling points. •
- ❖ They are insoluble in water.
- ❖ Most covalent compounds are soluble in non- polar solvents.
- ❖ Non-polar covalent compounds are non-electrolytes because they do not conduct electricity.

2012. Which of the following properties is true about covalent molecules?

- A. High melting but low boiling point
- C. Low melting but high boiling point
- B. High melting and low boiling point
- D. High electrical and thermal conductivity

Valence Shell Electron Pair Repulsion (VSEPR) Theory or model

- ❖ Model or procedure used to predict the shape/geometry/ of molecule or ions
- ❖ Used to convert two-dimensional Lewis structure (2D) into three-dimensional (3D) geometry.

Set of electrons or electron sets

- ❖ Defined as any number of electron pairs around a central atom.
- ❖ Consists of bond pair [single bond (–), double bond (=), triple bond (≡)], lone-pair(s) or a lone (single) electron.
- ❖ Ranges from 2 to 6.
- ❖ One set of electrons = one (single bond, double bond, triple bond, lone pair or single electron = one area of electron density.

Electron pair arrangement or geometry is:

- ❖ Defined by the sets of electrons (both bonding and nonbonding) around the central atom.
- ❖ Arrangement of electron pairs around central atom
- ❖ Determined from electron sets or pairs (bond pair and lone pair) = steric number

Pair or set of electrons	Electron pair arrangement
2 → (20)	Linear
3 → (21,30)	Trigonal Planar(TP)
4 → (22, 31,40)	Tetrahedral or tetrahedron (Td)
5 → (31,41,50)	Trigonal bi-pyramid (TBP)
6 → (42,51, 60)	Octahedral or octahedron (Oh)

Molecular shape or geometry is:

- ❖ Defined by the relative positions of the atomic nuclei.
- ❖ Arrangement of atoms in a molecule or compounds around central atom
- ❖ Determined from :
 - ✓ AX_mE_n (electron pairs), where A is the central atom, X is the terminal atom, E is lone-pair (nonbonding) electron sets, m is bond pair and n is non bonding pair
 - ✓ Number of bonding and non bonding pairs
- ❖ If there is no lone pair on the central atom electron pair arrangement and molecular shape are the same.

What is a steric number and what is the use of it?

Steric number is:

- Sum of bonded and unbounded electron pairs in the molecule (BP + LP)
- Can be used to predict shape of molecule
- Range from 2-6
- 1. Linear (20/23), 2. Bent or V-shape (21/22), 3. Trigonal planar (30), 4. Trigonal pyramid (31), 5. T-shape (32), 6. Tetrahedral (40), 7. Seesaw or distorted tetrahedral (41), 8. Square planar (42), 9. Trigonal bi-pyramid (50), 10. square pyramid (51) and
- 11. Octahedral (60).

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✓ The 1st number is **bond pair** and the 2nd number is **lone pair**

Decreasing/minimizing/diminishing order of Repulsion/distortion/ of valence shell electron pairs:

Lone pair vs lone pair > lone pair vs bonding pair > bonding pair vs bonding pair.

Causes or factors that affect distortion of molecules

A. Lone pairs (LP)

- Require more space than bonding pairs.
- Take/occupy/ more space than bonding pairs.
- Cause/exert/ more repulsive force and distortion in structure than bond pairs.

B. Pi- bonds (triple and double bonds)

Triple bonds are:

- Fatter than double bonds
- Double bonds are fatter than single bonds.
- Occupy more space than double and single bond.
- Cause more distortions than double or single bond
- Double bond cause more distortion than single bond.
- Angle involving Pi-bon are smaller than angle involving sigma bonds
- Electron-electron repulsion decreases in the order:

Triple bond > double bond > single bond.

C. Electro negativity difference between atoms

- Bond distorted toward the more electronegative atom.
 - ❖ Number of multiple or π - bond increase
 - ❖ Number of lone pair increase
 - ❖ Electron negativity difference increase
- Repulsion increase or distortion increase
Bond angle decrease

Molecular shape and molecular polarity

Molecular shape or geometry:

- Affects many properties of the molecule (physical and chemical properties)
 - ✓ Molecular polarity
 - ✓ Melting and Boiling points
 - ✓ Solubility
 - ✓ Reactivity.

Molecular polarity is:

- Created by molecules with a **net imbalance of charge**
- Determined by shape of molecule and polarity of bond
- **Cis- shape is polar molecule and trans-shape is non polar molecule.**

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Example1: Consider tetra chloromethane (CCl_4) and tri chloromethane (CHCl_3)

- ✓ Two AX_4 type molecules (they have tetrahedral shape with different polarities).

In CCl_4

- The surrounding atoms are all Cl atoms.
- Each C–Cl bond is polar
- The molecule is non-polar because the individual bond polarities counterbalance each other.

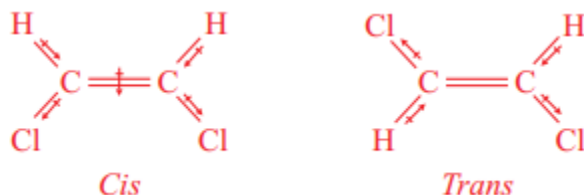
In CHCl_3

- ✓ H atom substitutes for one of the Cl atoms, disturbing the balance and giving chloroform a significant dipole moment.
- ✓ Polar molecule



Example 2: Consider the two constitutional isomers of di chloro ethane ($\text{C}_2\text{H}_2\text{Cl}_2$)

- They have the same molecular formula.
- They have different physical and chemical properties. VSEPR theory predicts that all the nuclei lie in the same plane with a trigonal planar molecular shape around each carbon atom.



Trans isomer	Cis-isomer
<ul style="list-style-type: none">➤ Non-polar molecule➤ Has no dipole moment ($\mu = 0 \text{ D}$) because the C – Cl bond polarities balance each other.	<ul style="list-style-type: none">➤ Polar molecule ($\mu = 1.90 \text{ D}$) because the bond dipoles partially reinforce each other➤ Dipole pointing between the Cl atoms

Example 3: Consider carbon dioxide and water.

- Both water and carbon dioxide are tri atomic molecules
- The large electro negativity difference between C and oxygen makes each C-O bond polar.

Carbon dioxide (CO_2) is:

- ✓ AX_2 type molecule with two sets of electrons around the central atom
- ✓ Its shape is linear

- ✓ Its bonds are directed 180° from each other, so there is no net dipole moment (μ); $\mu = 0$ D.

Water is:

- ✓ Has a significant dipole moment ($\mu = 1.87$ D).
- ✓ Due to the effect of the lone pairs of electrons on its shape, water is a polar molecule.
- ✓ In each O–H bond, electron density is pulled toward the more electronegative O atom, but the bond polarities do not counterbalance each other, because the water molecule is V-shaped.
- ✓ The bond polarities partially reinforce each other, and the oxygen end of the molecule is considerably more negative than the other end.

Exercise 3.10

1. In which situation/s the molecular shape and the electron-set arrangement is/are same?
2. Arrange the following AF_n species in order of increasing F–A–F bond angles. BF_3 , BeF_2 , CF_4 , NF_3 , OF_2
3. In the gas-phase, phosphorus pentachloride exists as separate molecules. In the solid phase, however, the compound is composed of alternating PCl_4^+ and PCl_6^- ions. What change(s) in molecular shape occur(s) as PCl_5 solidifies? How does the angle change?
4. For molecules of general formula AX_n (where $n > 2$), how do you determine if a molecule is polar?
5. For each of the following molecular formula (AX_mE_n), A is the central atom, X is the terminal atom, m is the number of terminal atoms bonded to the central atom; E is the lone-pair of electrons and n is the number of lone-pairs possessed by the central atom. The first item has been completed as an example. Fill the rest of the table after drawing in your note book.

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Answer:

1. When there is no lone pair of electron(s) on the central atom.
2. $\text{OF}_2 < \text{NF}_3 < \text{CF}_4 < \text{BF}_3 < \text{BeF}_2$
3. $\text{PCl}_5(\text{g})$ (AX_5 type) has a trigonal bipyramidal molecular shape. In the solid phase, PCl_4^+ (AX_4 type) is tetrahedral, and PCl_6^- (AX_6 type) is seesaw in shape.
4. By the vector sum of bond moments

Number of e-pairs	Molecular formula	Molecular geometry	Electron set geometry	Sketch	Examples
2	AX_2	linear	linear	X-A-X	Cl-Be-Cl
3	AX_2E AX_3				
4	AX_4 AX_3E AX_2E_2				
5	AX_5 AX_4E AX_3E_2 AX_2E_3				
6	AX_6 AX_5E AX_4E_2				

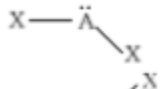
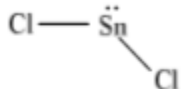
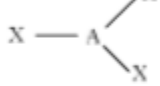
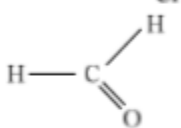
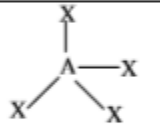
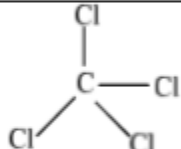
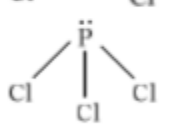
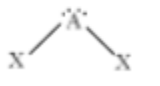
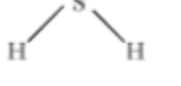
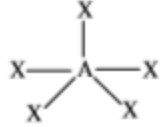
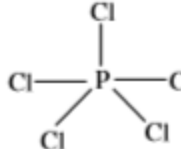
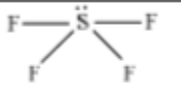
6. Determine the molecular shape and predict the bond angles (relative to the ideal angles) in

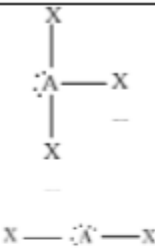
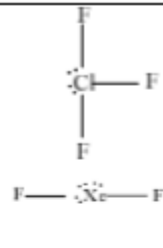
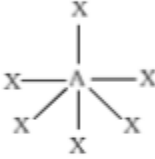
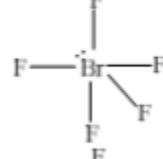
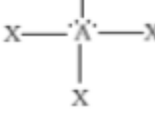
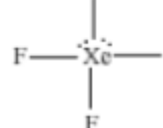
a SbF_5 b BrF_5

7. Predict the geometry of and bond angles (relative to the ideal angles) in the following molecules and ions.

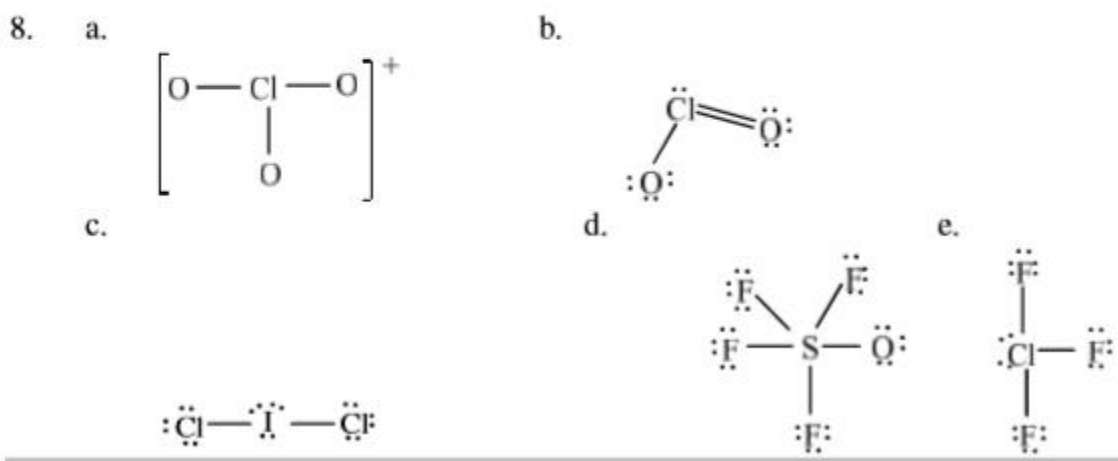
a ClO_3^+ b ClO_2^+ c ICl_2^- d SOF_4 e ClF_3

8. Draw molecular shapes for the molecules/ions in question number 8.

No of electron pairs	Molecular formula	Molecular geometry	Electron set geometry	Sketch	Example
2	AX ₂	Linear	Linear	X—A—X	Cl—Be—Cl
3	AX ₂ E	Angular	Trigonal planar		
	AX ₃	Trigonal Planar	Trigonal Planar		
4	AX ₄	Tetrahedral	Tetrahedral		
	AX ₃ E	Trigonal pyramidal	Tetrahedral		
	AX ₂ E ₂	Angular	Tetrahedral		
5	AX ₅	Trigonal bipyramidal	Trigonal bipyramidal		
	AX ₄ E	Seesaw	Trigonal bipyramidal		
	AX ₃ E ₂	T-shape	Trigonal bipyramidal		

	AX_2E_3	Linear	Trigonal bipyramidal		
6	AX_6	Octahedral	octahedral		
	AX_5E	Square pyramidal	octahedral		
	AX_4E_2	Square planar	octahedral		

6. a. Trigonal bipyramid b. square pyramidal
7. a. Trigonal planar (120°) b. Trigonal planar (120°)
- c. Angular (bent) (104°) d. Linear (180°)
- e. Trigonal bipyramidal ($90^\circ, 120^\circ$) f. T – shape (90°)



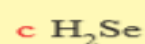
Bond Angle is:

- Formed by two surrounding atoms at vertex of central atom

Molecular shape	Bond Angle
Linear	180°
Trigonal planar	120°
Tetrahedral	109.5°
Trigonal pyramidal	107.5°
Vent or angular or V-shaped	104.5°
Trigonal bi-pyramid	$90^\circ, 180^\circ$
Oh, T-shape and square pyramidal	90°

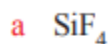
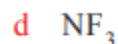
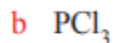
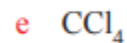
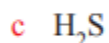
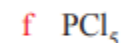
Exercise 3.9

Determine the molecular shape and ideal bond angles in

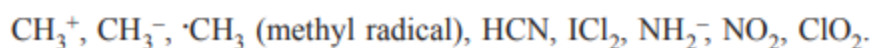


- a. Tetrahedral, 109° b. Linear, 180° c. Angular (bent), 104°

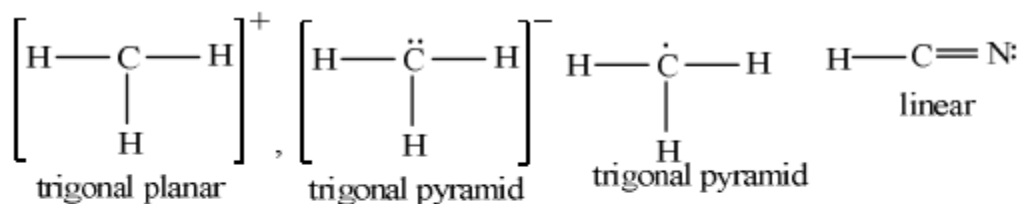
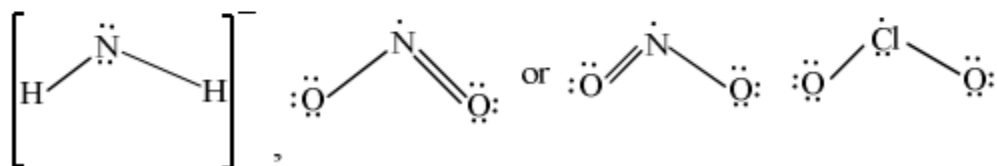
22. Predict the shape for the following molecules:



23. Draw Lewis structures and predict the shapes, giving approximate bond angles for:



22. a. Tetrahedral b. Trigonal pyramidal c. Angular (bent)
d. Trigonal pyramidal e. Tetrahedral f. Trigonal bipyramidal



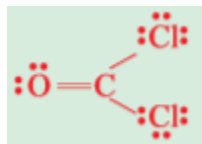
33. Based on the VB approach, predict the shape of a ClO_3^- ion

✓ AX_3E type – trigonal pyramid

✓

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2012. For the molecule phosgene shown below, why do the lone pairs and double bonds effects the bond angles?



- D. Lone pairs repel more due a greater charge density than bonding pairs.
- E. Lone pairs repel smaller than bonded pairs and multiple bonds have a higher charge density around the central atom.
- F. Both multiple bond and lone pairs occupy less space.
- G. Lone pairs occupy less space, but multiple bonds occupy more space than single bonds

2012. Based on the VSEPR model, what shape would you expect for the molecule $[I_3]^-$?

2012. Which of the following is **NOT** the correct geometric configuration for the given molecule?

- A. CH_4 , tetrahedral
- C. SF_6 , Octahedral
- B. NF_3 , trigonal planar
- D. BF_3 , trigonal Planar

2010. From CO_2 , H_2O , $BeCl_2$ and N_2O which have the same molecular geometry?

- A. CO_2 , H_2O and N_2O
- C. CO_2 , $BeCl_2$ and N_2O
- B. CO_2 and $BeCl_2$
- D. H_2O and N_2O

2009. Which of the following molecule is in correctly matched with the **electronic geometry** of central atom?

- A. CF_4 - Tetrahedral
- B. H_2O - Tetrahedral
- C. PCl_3 - Pyramidal
- D. $BeBr_2$ – linear

➤ CF_4 , H_2O and PCl_3 are surrounded by four set of electrons (40, 22 and 31) → have tetrahedral electronic geometry

- $BeBr_2$ is surrounded by two sets of electrons (20) → has linear electronic geometry

2008. Which of the following molecule is **NOT** probable shape of molecule?

- A. $GeCl_4$ - Tetrahedral
- B. HCN - Bent
- C. PCl_5 – Trigonal Bi-Pyramidal
- D. BrF_3 – T-shape

➤ $GeCl_4$ is surrounded by four set of electrons (40), → have tetrahedral electronic geometry and molecular geometry.

➤ HCN is surrounded by two set of electrons (20), → have linear electronic geometry and molecular geometry.

➤ PCl_5 is surrounded by five set of electrons (50), → have trigonal Bi-pyramidal electronic geometry and molecular geometry.

➤ BrF_3 is surrounded by five set of electrons (32), → have trigonal bi-pyramidal electronic geometry and T- shape geometry.

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2008. Which of the following molecules does **NOT** have a **trigonal bi-pyramidal electron geometry**?

A. SF_4

B. ClF_3

C. XeF_2

D. BrF_5

- SF_4 , ClF_3 and XeF_2 are surrounded by **five set of electrons (41, 32 and 23)** $\longrightarrow$ have **trigonal bi-pyramidal electronic geometry**

- BrF_5 is surrounded by **six sets of electrons (51)** $\longrightarrow$ has **octahedron electronic geometry**

2005. Which of the following molecule has a **trigonal bi-pyramidal structure**?

A. SF_4

B. IF_5

C. ICl_4^-

D. BrF_5

Molecule	Electron set	Molecular shape	Electron pair arrangement
SF_4	5 $\longrightarrow$ 41	Seesaw	Trigonal bi-pyramid
IF_5 / BrF_5	6 $\longrightarrow$ 51	Trigonal pyramidal bi-	Octahedral
ICl_4^-	6 $\longrightarrow$ 42	Square planar	

2008. Which of the following molecules **does not have a tetrahedral** central atom?

A. XeF_4

B. SiCl_4

C. BF_4^-

D. NH_4^+

2007. The molecular shape of the H_3O^+ ?

A. Linear

B. Tetrahedral

C. Bent

D. trigonal pyramidal

➤ H_3O^+ is surrounded by **four set of electrons(31)** $\longrightarrow$ have **tetrahedral electronic geometry and trigonal bi-pyramidal**

2006. Which molecule or ion does not have a **tetrahedral shape**?

A. SF_4

B. AlH_4^-

C. BF_4^-

D. SiCl_4

➤ SiCl_4 , BF_4^- , NH_4^+ , AlH_4^- are surrounded by **four set of electrons(40)**

$\longrightarrow$ Have tetrahedral **molecular and electronic geometry**

➤ XeF_4 is surrounded by **six sets of electrons(42)** $\longrightarrow$ has **octahedron electronic geometry and square planar molecular geometry**

➤ SF_4 is surrounded by **five sets of electrons(41)** $\longrightarrow$ has **trigonal bi-pyramidal electronic geometry and seesaw molecular geometry**

2003. What is the **molecular geometry** of O_3 ?

A. Trigonal Planar

B. Tetrahedral

C. Bent

D. Trigonal pyramidal

➤ O_3 is surrounded by **three set of electrons (21)** $\longrightarrow$ have **trigonal planar electronic geometry and bent molecular geometry**.

2003. NF_3 , BF_3 , ClF_3 all have formula of XF_3 . Which molecule(s) has (have) **trigonal pyramidal molecular geometry**?

A. NF_3 & BF_3

B. BF_3 & ClF_3

C. NF_3 & ClF_3

D. NF_3

➤ NF_3 is surrounded by **four set of electrons(31)** $\longrightarrow$ Has **trigonal molecular geometry and tetrahedral electronic geometry**

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- BF_3 is surrounded by three sets of electrons(30) $\longrightarrow$ has trigonal planar electronic and molecular geometry
- ClF_3 is surrounded by five sets of electrons(32) $\longrightarrow$ has trigonal bi-pyramidal electronic geometry and T-shape molecular geometry

1999. What is the shape of the TeF_5^- anion?

- A. Seesaw B. Square pyramid C. Trigonal pyramid D. Trigonal bi-pyramid

- TeF_5^- is surrounded by six set of electrons(51) $\longrightarrow$ Has square pyramid molecular geometry and octahedral electronic geometry

1999. In a molecule of ICl_3 , the arrangement of electron pair would be described as:

- A. Square planar B. Trigonal Bi-pyramid
C. Trigonal pyramid D. Octahedral

- ICl_3 is surrounded by five sets of electrons(32) $\longrightarrow$ has trigonal bi-pyramidal electronic geometry and T-shape molecular geometry

Questions related to Bond Angle

2010. Given the following AF_n species BF_3 , BeF_2 , CF_4 , NF_3 , and OF_2 .

What is the correct order of F-A-F bond angles?

Answer: - OF_2 (104.5°) < NF_3 (107.5°) < CF_4 (109.5°) < BF_3 (120°) < BeF_2 (180°)

2009/2010. What is (are) the F-S-F bond angle (s) in SF_6 ?

- A. 109.5° B. 120° C. $90^\circ/120^\circ$ D. 90° and 180°

2008. Which molecule has the largest bond angle?

- A. H_2O B. CO_2 C. NH_3 D. CH_4

2004. Which of the following species has the smallest H-X-H bond angle where X is central atom? A. OH_2 B. NH_3 C. BH_3 D. CH_4

1998. Which molecule has the largest bond angle?

- A. H_2O B. CO_2 C. NH_3 D. CH_4

Molecule	Electron sets	Molecular Geometry	Bond Angle
OH_2 , H_2O , OF_2 , NH_2^-	$22 \rightarrow 2+2 = 4$	Bent/V-shape/angular	104.5°
NF_3 , NH_3	$31 \rightarrow 3+1 = 4$	Trigonal Pyramid	107.5°
CF_4 , CH_4 , NH_4^+	$40 \rightarrow 4+0 = 4$	Tetrahedral	109.5°
BF_3 , BH_3	$30 \rightarrow 3+0 = 3$	Trigonal Planar	120°
BeF_2 , CO_2	$20 \rightarrow 2+0 = 2$	Linear	180°
SF_6	$60 \rightarrow 6+0 = 6$	Octahedral	90° or 180°

1998. What is the H-N-H bond angle in the three species NH_2^- , NH_3 and NH_4^+ respectively?

As lone pair increase = Bond angle decrease (NH_2^- (104.5°) < NH_3 (107.5°) < NH_4^+ (109.5°))

Home work (1996/1995)

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Intermolecular Force (IMF)

- ❖ There are two types of electrostatic forces. These are: intramolecular and intermolecular forces.

Electrostatic force	
Intramolecular force:	Intermolecular forces
<ul style="list-style-type: none"> ❖ Chemical bond (ionic, covalent or metallic) ❖ Force that exists within a particle (molecule or polyatomic ion) and ❖ Affects the chemical properties ❖ Stronger than IMF 	<ul style="list-style-type: none"> ❖ Attraction between ions and molecules ❖ Exist between particles ❖ Forces that holds molecules together (forces inside of molecule) ❖ Influence physical properties (mpt, Bpt) ❖ Responsible for condensation and sublimation of molecule ❖ Weaker than intramolecular force

There are Three Types of Intermolecular Force

1. London(Dispersion force)

- ❖ Exists in all atoms and non-polar molecules
- ❖ Weakest IMF

Example: Rn, Xe, Kr, Ar, Ne, He, H₂, O₂, N₂, F₂, Cl₂, Br₂, I₂, P₄, S₈, CO₂, SO₃, CCl₄, H₂,

- ❖ Strength of London force depends on the molecular size(mass) and number of electrons
- ❖ **Larger molecular size (weight) = larger number of electrons = larger number of atoms = (IMF, London force, boiling point, melting point) increase.**

Example1: F₂ < Cl₂ < Br₂ < I₂

Example 2: CH₄ < SiH₄ < SnH₄

2. Dipole-dipole force

- ❖ Found in polar molecule
- ❖ Stronger than London force.

Example: HCl, HBr, HF, NO, H₂S, SO₂, PH₃, NH₃, H₂O, CH₃OH

3. Hydrogen bonding

- ❖ Special type of DDF
- ❖ **Formed between hydrogen and FON**
- ❖ **Strongest IMF**

LF < DDF < HB < CB < IB

Example: HF, NH₃, H₂O, Carboxylic acid (CH₃COOH, CH₃CH₂COOH), lower alcohol (CH₃OH, CH₃CH₂OH) and Protein (DNA, RNA)

Example 1: NH₃ > SbH₃ > AsH₃ > PH₃

HB molecular mass

Example 2: H₂O > H₂Te > H₂Se > H₂S

HB molecular mass

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2012. Which of the following substances has London dispersion forces as its only intermolecular force? (No hydrogen bonding, no dipole- dipole forces).

A. CH₃OH

B. NH₃

C. H₂S

D. CH₄

HB

HB

DDF

LF

2004. Which of the following explains, why at room temperature I₂ is a solid, Br₂ is a liquid and Cl₂ is a gas?

A. Ionic bonding

B. Hybridization

C. Hydrogen bonding

D. London- dispersion force

❖ Molecular mass increase = Strength of IMF = State of substance change from solid to liquid and liquid to gas.

❖ I₂, Br₂ and Cl₂ are non polar molecule and contain London force

2003. Of Br₂, Ne, HCl and N₂, which is likely to have the largest intermolecular dispersion forces?

A. Br₂

B. Ne

C. HCl

D. N₂

❖ HCl has dipole –dipole force

❖ Br₂, Ne and N₂ are non polar molecule and contain London force

❖ Largest molecular mass = Largest IMF (LF)

❖ Br₂ (160), N₂(28) and Ne(20)

❖ Br₂ has largest London force because Br₂ has largest molecular mass or size

2001. Which of the hydrogen halide has the highest enthalpy of vaporization?

A. HCl

B. HBr

C. HI

D. HF

✓ IMF increase = heat of vaporation, Bpt and Mpt increase

✓ HCl<HBr<HI<HF

2000. Which type of bonding explains best high Bpt of water?

A. Vander Waals forces

B. Covalent Bonding

C. Polar ionic Bonding

D. Hydrogen Bonding

1999. Which type of molecular attraction accounts for high Bpt of water ?

Compound	H ₂ O	H ₂ S	H ₂ Se	H ₂ Te
Bpt	100°C	-60.7 °C	-4.5 °C	-2.2 °C

A. Vander Waals forces

B. Molecular

C. Ion-ion

D. Hydrogen Bonding

✓ Hydrogen bond is responsible for high Bpt of water.

1999. Which type of IMF exists in non polar molecule?

A. Dipole- Dipole forces

B. Ion –dipole force

C. London force

D. Hydrogen Bonding

Intermolecular force	Strength	Polarity
Hydrogen Bonding	The Strongest	Polar
Dipole-dipole Forces	Moderate	Polar
London Dispersion Forces	The Weakest	Non-polar

Activity 3.16

1. Why does a polar liquid generally have a higher normal boiling point than a nonpolar liquid of the same molecular mass?
2. State the principal reasons why CH_4 is a gas at room temperature whereas H_2O is a liquid.
3. Only one of these substances is a solid at STP: $\text{C}_6\text{H}_5\text{COOH}$, $\text{CH}_3(\text{CH}_2)_8\text{CH}_3$, CH_3OH , $(\text{CH}_3\text{CH}_2)_2\text{O}$. Which do you think it is and why?
4. Only one of these substances is a gas at STP: NI_3 , BF_3 , PCl_3 , and CH_3COOH . Which do you think it is and why?

Answer:

1. When we compare and contrast physical properties like boiling point and solubility, we usually consider comparable masses of substances under consideration.

The intermolecular force in polar liquids is the dipole – dipole force (stronger), and in non-polar liquids it is London forces (weakest). It is easier to boil non-polar liquids than the polar ones.
2. The more polar a molecule – that is, the greater the dipole moment – the more pronounced is the effect of the dipole – dipole forces on physical properties. Hydrogen bonds (strongest) hold water molecules in a rigid structure keeping them very close to each other, whereas the weak London (dispersion) forces in methane cannot bring distant particles close to each other. This is one of the reasons that the particles of gases are far apart.
3. Benzoic acid and methanol are different from the other two-diethyl ether and decane. There is hydrogen bond in both benzoic acid and methanol, but there are two possible ends for hydrogen bonding in benzoic acid. The larger mass and the two possible ends in this acid make it solid at STP. Benzoic acid is a white crystalline solid. The rest are liquids under these conditions.
4. PCl_3

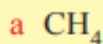
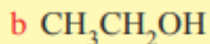
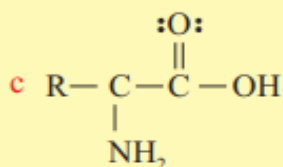
Activity 3.17

1. Name some intermolecular forces.
2. Cis-1,2-dichloroethene shows higher boiling point than trans-1,2 dichloroethene. Explain.
3. Which type of compounds exhibit hydrogen bonding?
4. Name the factors on which the strength of dispersion forces of the particles depend?
5. 'Ice is lighter than water.' Comment

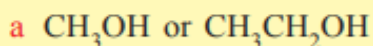
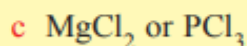
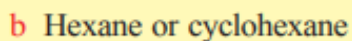
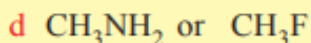
1. Dipole – dipole, hydrogen bond, London (dispersion) forces, etc.
2. The trans-isomer is non-polar, and the intermolecular force governing these types of molecules is the weak dispersion forces, whereas the cis-isomer is polar. The intermolecular force existing in polar covalent molecules is the strong dipole forces, which keeps the particles of the substance very close to each other. Therefore it exhibits high boiling points.
3. Compounds in which hydrogen is bonded to highly electronegative and smaller sized atoms such as nitrogen, oxygen, and fluorine form hydrogen bonding. These atoms exert a strong attraction for the bonding electron pair. The effect is very much like that of a bare proton at the end of the bond. It is represented as $X-H\cdots Y$, where X is the atom to which the hydrogen is actually normally bonded, and Y is the atom to which the hydrogen is bonded. It has been found experimentally that hydrogen bonds of significant strength are formed when the atoms X and Y are either N, F, or O. If either X or Y is another element (e.g., Cl), much weaker hydrogen bonds are formed.
4. Polarizability of large molecules, molecular mass and molecular geometry are the essential factors to consider.
5. The hydrogen bond in ice holds the water molecules in a rigid but open structure and more volume is occupied. As ice melts, some of the hydrogen bonds are broken. Water molecules move into some of the holes in the rigid structure of ice. This means that liquid water at 0°C is more dense than ice.

Exercise 3.11

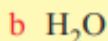
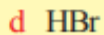
1. In which of the following substances do hydrogen bonds occur? Explain with the help of diagrams.



2. Identify the dominant intermolecular force that is present in each of the following substance and select the substance with the higher boiling point in each pair:



3. Which type of intermolecular force is dominant in the following substances?



4. Compare intermolecular forces with that of intramolecular bonding.

Answer:

1. In b and c.

2. a. ethanol (hydrogen bonding)

- b. hexane (Dispersion forces)

- c. magnesium chloride (ionic)

- d. methyl amine (hydrogen bonding)

3. a. dipole – dipole

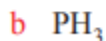
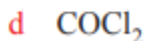
- b. Hydrogen bonding

- c. London

- e. Dipole – dipole

4. Intramolecular forces are chemical bonds which are stronger than any of the intermolecular forces.

12. In which of the following compounds would inter-molecular hydrogen-bonding occur?



13. Which of the following molecules would you expect to have the highest boiling point?
- a $\text{CH}_3\text{--CH}_2\text{--CH}_2\text{--CH}_2\text{--CH}_2\text{--CH}_3$
 - b $\text{CH}_3\text{--CH}(\text{CH}_3)\text{--CH}_2\text{--CH}_2\text{--CH}_3$
 - c $\text{CH}_3\text{--CH}_2\text{--CH}(\text{CH}_3)\text{--CH}_2\text{--CH}_3$
 - d $\text{CH}_3\text{--C}(\text{CH}_3)_2\text{--CH}_2\text{--CH}_3$
14. Which of the following molecules possesses the strongest forces between the molecules?
- a N_2
 - b $\text{CH}_3\text{--CH}_3$
 - c H_2CO
 - d CH_4
15. Which one of the following usually produces the weakest interaction between particles of similar molar mass?
- a Van der Waals forces
 - b Dipole-dipole forces
 - c Covalent bonds
 - d Hydrogen-bonding
16. The hydrogen bonding is not significant in:
- a DNA
 - b protein
 - c ice
 - d polystyrene
12. C 13.A 14. C 15. A 16.D
29. Explain the following in terms of the intermolecular forces that exist in them:
- a Water is a liquid at room temperature and atmospheric pressure, but hydrogen sulphide is a gas.
 - b At room temperature and atmospheric pressure chlorine is a gas, bromine is a liquid and iodine is a solid.
 - c Pentan-1-ol boils at 137°C , whereas pentan-3-ol boils at 116°C .
 - d The boiling point of sulphur dioxide is 24°C higher than that of chlorine.

Answer:

29. a. Hydrogen bond in water
- b. London/dispersion forces/ increases as mass increases
- c. London/dispersion forces/ in branched chain is less than that of straight chains
- d. Dipole-dipole forces in SO_2

Metallic Bonding is:

- ❖ Bonding in metal atoms (Ni, Mg, Pt, Pd, Cu, K, Fe, Al)
- ❖ Electrostatic force between metal cations and delocalized electrons.
- ❖ Interaction between metal nuclei and delocalized electrons.
- ❖ Force of attraction which binds together the positive metal ions or Kernels with the electrons
- ❖ Imagined as sea of electrons in which positive metal ions are embedded.

Delocalized electrons are:

- ✓ Called as conduction electrons
- ✓ Mobile electrons
- ✓ Electrons which are not confined (attached) to a particular atom
- ✓ Not associated to one atom.
- ✓ Bonding electrons that spread over several or all atoms..

Metals nuclei are:

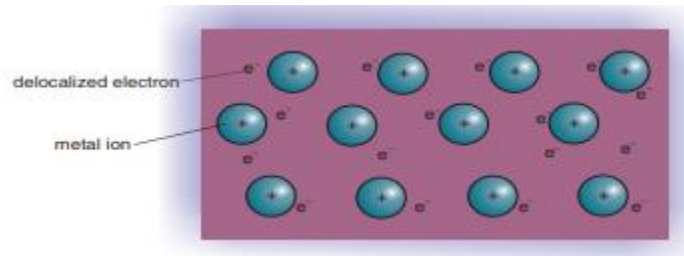
- ✓ Positive metal ions (Metal with core electrons)
- ✓ Metal nuclei with core electrons
- ✓ Positive metal ions are called Kernels.

The strength of the metallic bond depends on the:

1. Number of valence electrons in the delocalised 'sea' of electrons.
 2. Packing arrangement of the metal atoms.
 3. Size of cations
- ❖ More delocalised electrons, more closely packed the atoms and smaller size of cations are results in a stronger bond and a higher melting point.
 - ❖ Strength of metallic bonding decrease down a group and increase across a period.
 - ❖ Metallic character increase down a group and decrease across a period.

Electron-Sea Model is:

- ❖ Model which represents the electronic structures of metals
- ❖ All metal atoms contribute their valence electrons to form "sea" of electrons
- ❖ The metal ions (the nuclei with their core electrons) are submerged in the electron sea



Properties of Metals

Metals are:

- ❖ **Malleable** (hammered into thin sheets)
- ❖ **Ductile** (drawn into tubes or wires)
- ❖ **Durable, hard, Strong and opaque**
- ❖ **Good conductors of heat and electricity** because of their mobile electrons.
- ❖ **Sonorous** (have deep sound when hit), **Shiny, Lustrous** and **very cool**
- ❖ **Solids with high melting and much higher boiling points;** because the atoms of metals have strong attractive forces between them and much energy is required to overcome this force.
- ❖ **Metals have a wide range of melting points,** (from - 39°C (mercury) to 3410°C (tungsten).
- ❖ **Metals have high MPt, Bpt, density and conductivity of heat and electricity.**

Activity 3.18

1. Two characteristics that lead to an element being classified as a metal.
2. Type of bond existing in solid copper.
3. What is delocalisation of electrons?
4. Electron sea model.
5. Even though so much energy is required to form a metal cation with a 2+ charge, the alkaline earth metals form halides with general formula MX_2 rather than MX . Explain.
 1. **Metals are very reactive and form compounds with other elements quite easily.**
Sodium (Na) and potassium (K) are some of the most reactive metals.

Some of the physical properties which characterize metals are:
 - a. **They have high electrical and thermal conductivity.**
 - b. **They have bright lustrous appearance.**
 - c. **They are malleable and ductile.**
 - d. **They are very cool and shiny.**
 2. **Metallic bond.**
 3. **Electrons which are not confined to a particular atom (mobile electrons).**
 4. **The model which represents the electronic structure of metals.**
 5. **The alkaline earth metals have the valence shell electron configuration ns^2 .**
Therefore, their atoms show a tendency to give up two electrons to form 2+ ions.
which makes them isoelectronic with the nearest noble gas.

Activity 3.19

1. Take a piece of metal and a piece of wood of the same size.
2. Let these remain on a table so that both attain room temperature.
3. Place your left hand on the metal and right one on the wood.

What do you observe? Which hand feels colder? Why?

Answer:

- ✓ Your left hand feels colder because of the high heat conductivity of metals. Heat flows from your hand to the metal object.

Activity 3.10

1. Use the periodic table to arrange the following in order of increasing melting point K, Br₂, Mg and O₂. Explain the reason for the order you choose.
2. List four physical characteristics of a solid metal.
3. Briefly account for the following relative values:
 - a. Lithium boils at 1317°C and melts at 179°C. The boiling point is about 1138°C higher than its melting point.
 - b. The melting point of Li and Be are 180°C and 1287°C, respectively.
 - c. The melting point of sodium is 89°C, where as that of potassium is 63°C.
4. For each of the following pairs, which element will have the greater metallic character:
 - a. Li or Be
 - b. Li or Na
 - c. Sn or Pb
 - d. B or Al? Explain briefly.

1. $O_2 < Br_2 < K < Mg$ because oxygen and bromine are non-metals with covalent bonds. The dominant intermolecular force governing these molecules is the London (dispersion) force. The strength of London forces increases with molecular mass. Br_2 – liquid at room temperature is larger in mass than O_2 – gas at room temperature. Correspondingly, K and Mg are metals with 1+ and 2+ oxidation states which will contribute 1 and 2 electrons, respectively, for metallic bonding. The higher the number of electrons contributed, the stronger is the metallic bond and the higher are the melting and boiling points.
2.
 - a. They have high electrical and thermal conductivity.
 - a. They have bright lustrous appearance.
 - b. They are malleable and ductile.
 - c. They are hard.
3.
 - a. The boiling point is higher due to the large amount of energy necessary to separate the metal ions from each other in the electron “sea”.
 - b. Be has two valence electrons per atom, compared with Li which has one. The metallic bond is therefore stronger for Be.
 - c. K is a larger atom than Na. Therefore its electrons are held more loosely, and thus its metallic bond strength is weaker.
4.
 - a. Li is more metallic than Be. Metallic character decreases from left to right across a period.
 - b. Na is more metallic. Metallic character increases down a group.
 - c. Pb is more metallic for the same reason given in b.
 - d. Al is more metallic.

2006. Why are metals soft and malleable ?

- A. Because they are very shiny.
- B. Because they experience electrostatic repulsion
- C. Because of the presence of mobile electrons.
- D. Because the metals cations can slip over each other.
✓ Metals are malleable and ductile because atoms of metallic lattice are capable of slipping with respect to one other.

2004. Which of the following elements has the highest Melting point ?

- A. Iodine
 - B. Tungsten
 - C. Mercury
 - D. Bromine
- ✓ Metals have high melting point, boiling point and density.

Chemistry Handout For Grade -11

2002. Which of the following is **not true** of metallic bonding ?

- A. It gives rise to excellent electrical conductivity
- B. Electrons are free to move throughout the structure.
- C. The strength of metallic bonding increase down a group
- D. The strength of metallic bonding affects the melting and boiling point of metals.
 - ✓ Atomic size increase = metallic character increase = Metallic bonding decrease (down a group)

Q. How can you distinguish a metallic bond from an ionic bond or a covalent bond ?

- ✓ Metallic bonding is bonding within the atoms of a metal
- ✓ It is the hybrid form of ionic and covalent bonding.

CHEMICAL BONDING THEORIES

- ❖ The two modern chemical bonding theories are the **Valence Bond Theory (VBT)** and the **Molecular Orbital Theory (MOT)**.

2012. Which of the following is a basic chemical bonding theory ?

- A. Molecular orbital theory
- B. Covalent bonding theory
- C. Ionic bonding theory
- D. Valence Shell Electron Repulsion (VSEPR) Theory

A. Valence Bond Theory (VBT)

- ❖ Also called localized electron model
- ❖ Explains formation of covalent bonds in terms of overlap of half-filled atomic orbitals.
- ❖ Covalent bond is formed by overlapping of orbitals.
- ❖ Covalent bond is formed when a half-filled orbital of one atom overlaps with a half-filled of another.
 - ✓ Each bond involves two atoms
- ❖ Treats electrons as localized between pairs of atoms
- ❖ Leads to concept of hybridization.

2012. Which statements is NOT correct about the valence bond Theory ?

- A. According to the valence bond theory, hybrid orbitals are formed by overlapping of atomic orbitals.
- B. According to valence bond theory, molecular orbitals are formed by combining atomic orbitals.
- C. Valence bond theory treats electrons as localized between two atoms
- D. According to valence bond theory, each bond only involves two atoms.

Overlap

- ✓ Sharing a common region in the space
(Two separate atom orbitals occupy the same space).
- ❖ The overlap region, which is between the nuclei, is occupied by a pair of electrons.
- ❖ The strength of the bond between any two atoms depends on the extent of the overlap between the two orbitals.

Chemistry Handout For Grade -11

- ❖ Number of overlapping increase = strength of covalent bond increase.
- ❖ Greater overlap = stronger bond

Types of overlap : Head to head and side to side overlap

Overlap	
Head to head overlap	
<ul style="list-style-type: none"> ❖ Also Head on, end-to-end, end on or linear overlap and ❖ Include s – s, s – p and p-p overlap along the nuclear axis ❖ Results only in sigma bonds. ❖ Strong bond ❖ Concentrated electron density along bond axis 	<ul style="list-style-type: none"> ❖ Also Side-on, side-way, lateral or parallel overlap and ❖ Include only p-p overlap ❖ Form Pi- overlap ❖ Weak bond ❖ Concentrated electron density above and below bond axis

Bond	
Sigma(σ) bond:	Pi(π) - bond
<ul style="list-style-type: none"> ❖ Formed by head to head overlap ❖ Presents in the molecule that contains single bond. <p>Example: F₂, Cl₂, Br₂, I₂, HBr, HI, HF, HCl, H₂O, CH₄, CCl₄, NH₄⁺</p> <ul style="list-style-type: none"> ❖ Exists on every bond (single bond, double bond and triple bond) ❖ Has highest electron density along the bond axis ❖ All single(sigma) bonds have their electron density concentrated along the bond axis ❖ Bond axis is an imaginary line joining the nuclei in between the bonded nuclei . 	<ul style="list-style-type: none"> ❖ Formed by side-side overlap ❖ Exists in the molecule that contains multiple bond(double bond and multiple bond) <p>Example: O₂, N₂, CO₂, COS, HI, SO₂, C₂H₄, C₂H₂</p> <ul style="list-style-type: none"> ❖ Exists on the double bond and triple bond) ❖ It has two regions of electron density. One above and one below the sigma bond axis. ❖ The double-bond increases electron density between the nuclei.

Note:

- A single-bond always consists of one σ bond.
- A double-bond always consists of one σ and one π bond.
- A triple bond consists of one σ and two π bonds.
- The extent of overlap influences bond strength, however, many factors, such as lone - pair repulsions, bond polarities, and other electrostatic contributions affect overlap and the relative strength of σ and π bonds between other pairs of atoms.
- Double bond is about twice as strong as a single bond, and a triple bond is about three times as strong as a single bond

2012. Which types of intermolecular forces are present in molecular crystalline solids?

- Ionic, dispersion, dipole-dipole, hydrogen bonding
- Dispersion, dipole-dipole, hydrogen bonding

- C. Covalent, dispersion, dipole-dipole, hydrogen bonding
- D. **Metallic**, dispersion, dipole-dipole, hydrogen bonding